SACE ONE - AUSTRALIAN CURRICULUM

CHEMISTRY

WORKBOOK
THIRD EDITION

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Preface

As you progress through this workbook, you will become familiar with the core topics of the SACE Stage 1 Chemistry course. The workbook follows the SACE Subject Outline for this new Australian Curriculum subject. We chose to present these topics in this order, though you may encounter them in a different sequence with your teacher. Whichever way you read it, this workbook covers the entire year's material.

You can use this book in class, or at home to supplement your studies.

Not only does this workbook present key theoretical knowledge, but it also introduces examples that show science as a human endeavour. Questions and exam-style topic tests will familiarise you with how chemists work, and how the subject of chemistry is assessed.

Each author of this book is also a teacher, so we know what it's like to be in the classroom. We know that students prefer worked solutions, so you can learn the steps from the answers. The worked solutions for each question can be found in the final section of the book, organised by chapter sequence. We have also allocated marks to each question, and mapped them against the SACE Performance Standards. This combined approach means you will learn how to answer questions, and how those questions may be marked.

All students and teachers should refer to the SACE website – **www.sace.sa.edu.au** – for the most up-to-date version of the Subject Outline and Performance Standards.

We trust that you will find this book a useful and engaging introduction to senior chemistry.

All the best for your SACE Stage 1 studies.

Kate Morgante
Luke Starczak
Stephen Tulip

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Topic 1: Materials and their atoms

1.1 Properties and uses of materials

Science understanding

The uses of materials are related to their properties, including solubility, thermal and electrical conductivities, melting point, and boiling point.

Chemistry is, fundamentally, the study of matter and its properties. **Matter** has mass and takes up space (e.g., a stone, air, your eyeball, the Earth). Matter is distinct from **energy**, which does not have mass or take up space (e.g., light, sound, motion). All matter is composed of particles: small 'units' of a material that cannot be broken apart any further without changing the identity of the substance. Most materials on Earth are made up of atoms, ions, molecules, or a mixture of these three.

Materials are classified according to their **properties**, which are characteristics that define the substance. The properties of a material relate directly to the nature and arrangement of the particles comprising the material. Properties of materials can be classified as either physical properties or chemical properties.

Physical Properties

Physical properties are traits of a material that can be observed directly without any chemical reaction occurring, e.g., colour, length, density, melting point. Physical properties can be either extensive or intensive.

Extensive physical properties change in proportion with the amount of the material. For example, mass is an extensive physical property because if the amount of material is doubled, its mass will also double. Other extensive physical properties include volume and number of particles.

Intensive physical properties do not change when the amount of the material changes. Density is one example of an intensive physical property: the density of water is approximately 1 g/mL, and if the amount of water is doubled the density remains 1 g/mL (i.e., it does not become 2 g/mL). Other intensive physical properties include colour, lustre, temperature, malleability, ductility, viscosity, and solubility. Because intensive physical properties do not depend on the amount of substance, they are useful for identifying materials and separating mixtures.

Solubility

When a material dissolves in a liquid to create a homogenous mixture, the dissolved material is called the **solute**, the liquid is called the **solvent**, and the mixture is called a **solution**. Solubility measures the amount of a material that can be dissolved in a given liquid at a given temperature. For example, at 25 °C, about 200 g of sugar can be dissolved in 100 mL of water (equivalent to 1 decilitre), so the solubility of sugar is 200 g/dL at 25 °C.

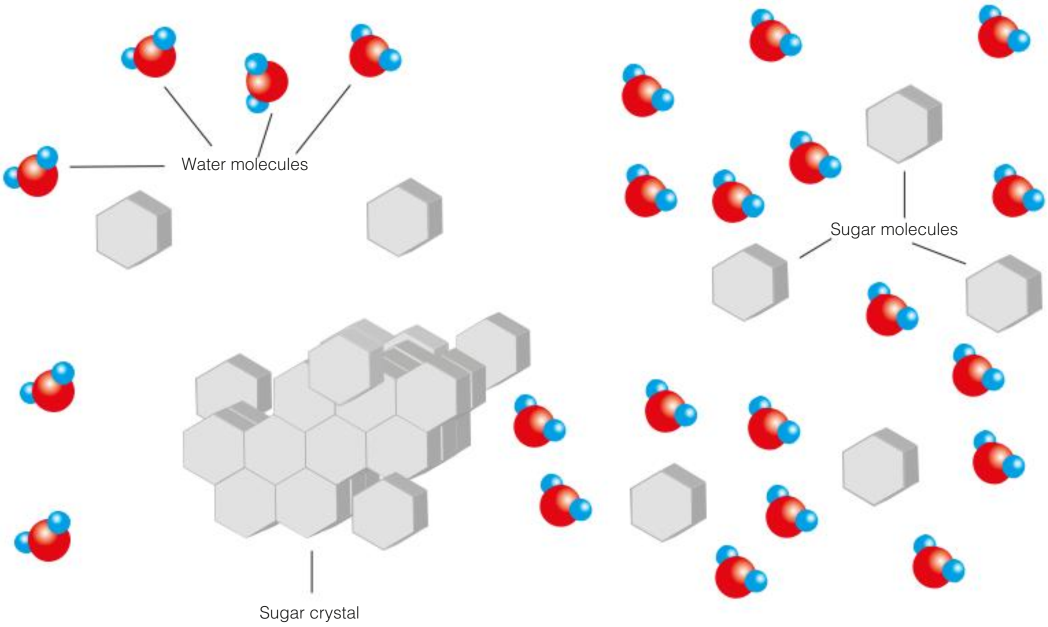

Figure 1.01: Sugar dissolving in water

For building or manufacturing, it is usually undesirable to use solid materials that have any significant solubility in water. Containers or packaging that will come into contact with food or beverages should ideally be totally **insoluble** so that the material does not contaminate the product.

Thermal Conductivity

The particles in substances are constantly in motion, and this motion of particles is experienced as heat. The amount of heat energy a material contains is an extensive physical property corresponding to the total kinetic energy of the particles in the material. Temperature represents the average kinetic energy and is an intensive physical property.

Thermal conductivity is a measure of the ability of a material to transmit heat through conduction. Conduction occurs when particles, in their normal course of motion due to their heat energy, collide with other particles and transfer some of that kinetic energy; this is analogous to a billiard ball striking another billiard ball and transferring some momentum. Materials with high thermal conductivity transfer heat rapidly and are called **thermal conductors**, while materials with low thermal conductivity transfer heat more slowly. Materials with very low thermal conductivity are called **thermal insulators**. Many applications require either maximised or minimised heat transfer, also called **thermal exchange**.

In general, metals are excellent thermal conductors, while other materials are less so.

Example

Hot beverages are generally not served in glasses because glass has a relatively high thermal conductivity. This allows heat to be rapidly transferred from the beverage through the glass to your hand, which might cause a burn. However, a thermos prevents thermal exchange by using materials with low thermal conductivity (e.g., plastic) and by separating the beverage from the exterior of the thermos with a vacuum chamber. A pure vacuum contains no matter, so it does not conduct heat at all.

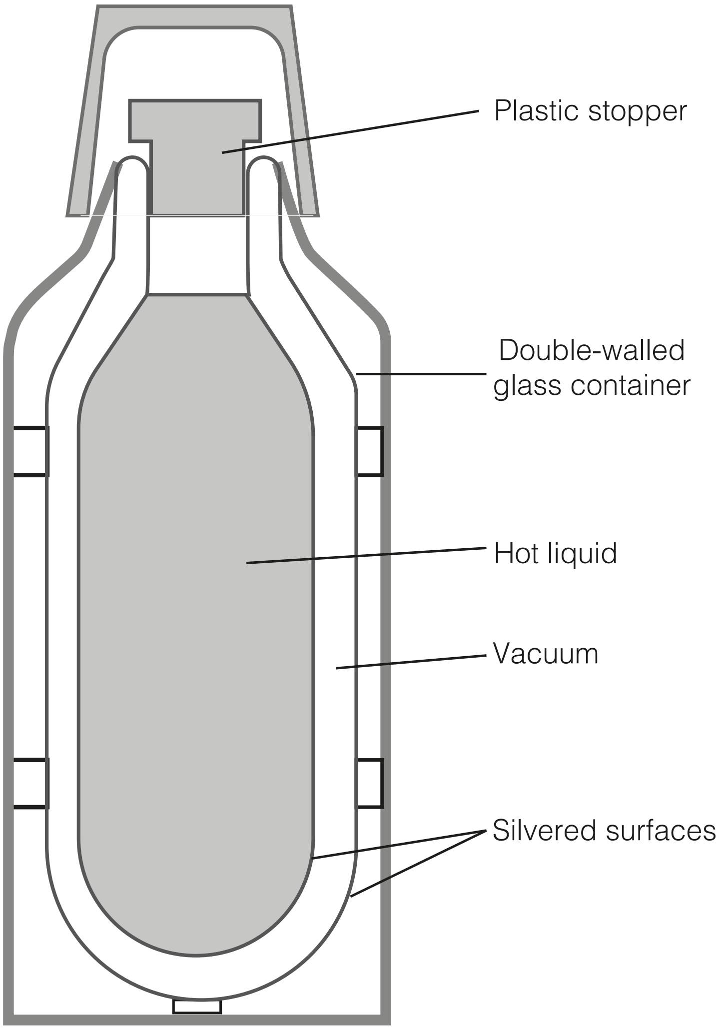

Figure 1.02: Diagram of a thermos

Science as a human endeavour

Window Insulation

Many Australian homes were built during a time when electricity was inexpensive and the danger posed by climate change was underappreciated. Little thought was given to thermal insulation, and many household windows are constructed of a single pane of glass in an aluminium frame.

A rising interest in home energy efficiency led by the environmentally and economically conscious has caused a surge in double-glazed window retrofitting. Double-glazed windows use two panes of glass with an air gap to reduce thermal exchange by 60–80%, keeping the heat in during winter, and out during summer.

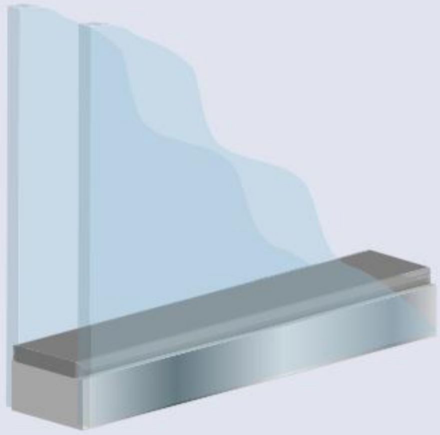

Figure 1.03: Cut-away diagram of double-glazed windowpane

Electrical Conductivity

The particles in a material may be neutral or may carry an electric charge. The motion of electric charge is called **current electricity** (or just 'electricity'), and the ability of a material to carry electricity through the movement of its charged particles is called **electrical conductivity**.

Some materials contain charged particles that are free to move with little resistance; these are good **electrical conductors**. Metals (e.g., copper) are good conductors because they have delocalised electrons which are free to move. As shown in Figure 1.04, dissolved ions (e.g., sodium chloride dissolved in seawater) will also move toward electrodes with opposite charges. The motion of these ions represents a flow of electricity, making solutions of ionic substances good conductors. Pure water contains no free charges and does not conduct electricity.

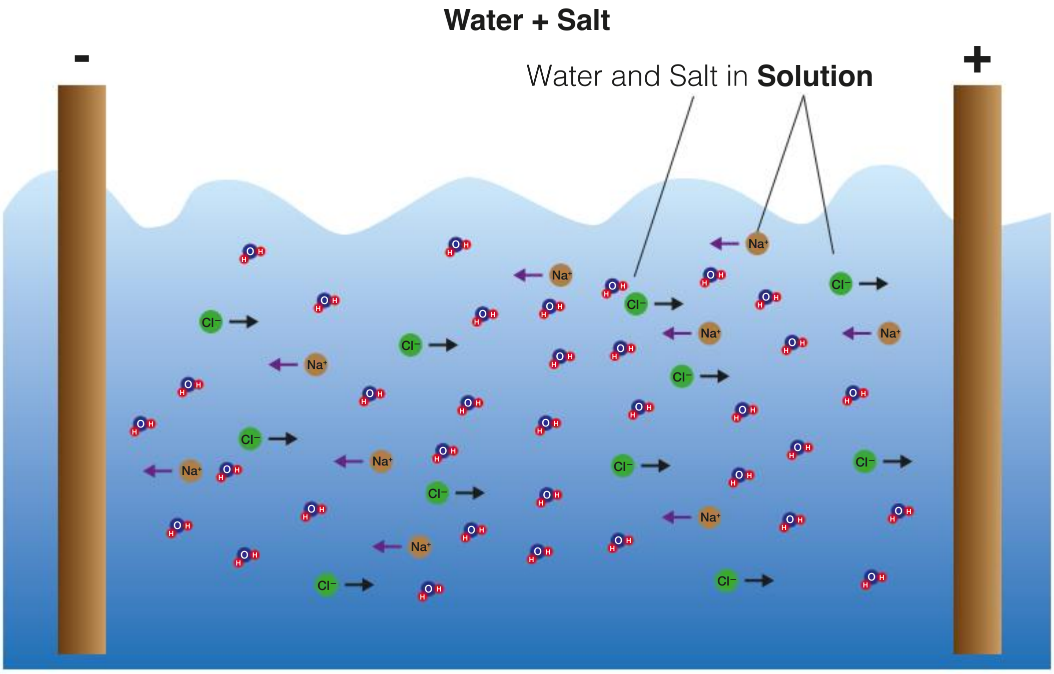

Figure 1.04: Ionic substances dissolved in water can carry electricity

For applications involving transmitting electricity, such as electrical wiring, high conductivity is desirable.

Other substances can carry electricity, but the charged particles encounter more resistance as they do so, and this causes the material to heat up as electricity is transmitted. These are called '**high-resistance conductors**', and usually consist of metal **alloys** that have been specifically designed with high resistance and high melting points in order to reliably produce heat when operated (e.g., the heating element in an electric oven). One common example is nichrome, an alloy of nickel and chrome.

Many materials cannot conduct electricity at all, either because they contain charged particles that are not free to move or no charged particles at all; these are **electrical insulators**. Examples include gases (nitrogen, oxygen, etc) and most non-metallic solids (glass, rubber, plastic, dry wood, etc.). Electrical insulators are used to control and limit electrical current. For example, electrical wiring is coated with plastic to prevent electricity from being conducted into objects the wire is touching.

Changes of State

Changes of state between solid, liquid, and gas phases are physical changes because they do not change the chemical identity of the material. When water freezes into ice, no chemical change has occurred: both ice and water are composed of molecules of water (H_2O). The melting point and boiling point of a substance are therefore intensive physical properties.

Melting Point

Solid materials are composed of particles that are bonded together in a rigid lattice. The position of the particles, relative to other particles in the same material, doesn't change over time. However, the particles do vibrate with heat energy, and this determines the temperature of the solid material. If sufficient thermal energy is transferred to the material, its temperature becomes high enough that the particles can break out of the lattice and begin to move relative to one another. The solid becomes a liquid; this process is called **melting**.

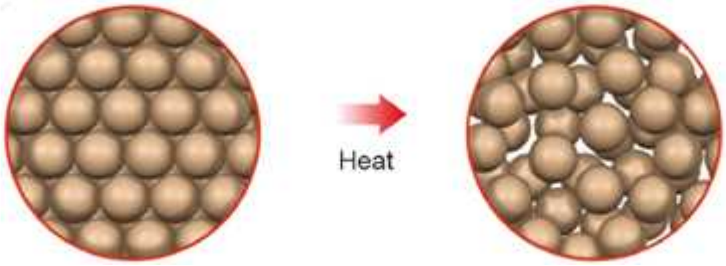

Figure 1.05: The change in behaviour of particles at the melting point

The temperature at which the transition from solid to liquid occurs is called the **melting point** and is characteristic of materials at defined pressures. For example, water has a melting point of 0 °C at atmospheric pressure, while sodium chloride has a melting point of 801 °C. The reverse transition, from liquid to solid, occurs at precisely the same temperature, so the "freezing point" of a material is always identical to the melting point.

Materials with high melting points are used when an object must withstand extremely high temperatures without melting. For example, kilns are usually made from brick rather than metal, as most metals have melting points lower than the temperature necessary for heating clay and other ceramics to set them.

Boiling Point

Liquid materials are composed of particles that are held together less tightly than those in solids. Particles in liquids can move relative to one another, but the particles remain bound together as a whole. If sufficient thermal energy is transferred to the material, its temperature becomes high enough that each particle can overcome the attractive forces and break away from the other particles altogether. The liquid becomes a gas; this process is called **boiling**.

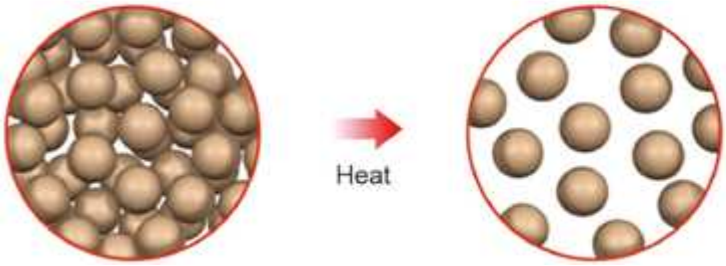

Figure 1.06: The change in behaviour of particles at the boiling point

The temperature at which the transition from liquid to gas occurs is called the **boiling point** and is also characteristic of materials at defined pressures. For example, water has a boiling point of 100 °C, while ethanol (alcohol) has a boiling point of 78 °C. The reverse transition, from gas to liquid, occurs at precisely the same temperature, so the "condensation point" of a material is always identical to the boiling point.

Liquids with high boiling points are used as coolants because this reduces the risk that the liquid will evaporate into a gas, which could cause an explosion. For example, commercially available coolant mixtures for vehicles usually have a boiling point above 120 °C, which is safer than using water.

Chemical Properties

A **chemical property** describes the way a material can undergo chemical changes. Observing a chemical property always requires a chemical reaction to occur, and the original substance must change its identity in that process.

Example

Acidity is a chemical property: it is a measure of the ability of a substance to donate a hydrogen cation (H^+). Sulfuric acid (H_2SO_4) is much more acidic than water (H_2O) because the hydrogen atoms in sulfuric acid can more easily take part in reactions. Sulfuric acid can rapidly dissolve metals because of its high acidity. However, this can only be observed by allowing the sulfuric acid to react with the metal, and this process consumes and transforms the sulfuric acid itself.

Chemical properties are the focus of Topics 5 and 6.

Questions

1. Most substances on Earth exist as either solids, liquids, or gases.

(a) In the three boxes below, draw a representation of the arrangement of particles in each of the three states of matter.

solid	liquid	gas

(3 marks) **KA1**

(b) With reference to the particles comprising the liquid, explain why a liquid has a defined volume but not a defined shape.

(2 marks) **KA1**

(c) Define the term ‘boiling point’.

(1 mark) **KA1**

2. For each of the following items, identify the property or properties of the material that make it suited to that use.

Material	Use	Property/Properties
aluminium	cookware	high thermal conduction high melting point
plastic	cookware handles	
wool	gloves	
copper	wiring	
nichrome	heating elements	
Kevlar	bulletproof vests	
gelatin	medicated gel capsules	

(6 marks) **KA1**

3. Fibreglass insulation used in walls and roof spaces consists of extremely fine glass fibres with pockets of air trapped between them.

Figure 1.07: Fibreglass insulation

(a) Glass is a moderately good thermal conductor. Describe the process of thermal conduction.

(2 marks) **KA1**

(b) Gases are very poor thermal conductors. Explain how the arrangement of particles in a gas reduces the rate of thermal exchange.

(2 marks) **KA1**

(c) Explain how fibreglass can be an excellent thermal insulator despite being made from glass.

(2 marks) **KA2**

Science understanding

Nanomaterials are substances that contain particles in the size range 1–100 nm.

- Suggest uses of materials, including nanomaterials, given their properties and vice versa.

Solubility, conductivity, and melting and boiling points are examples of **bulk properties**, which are properties caused by many particles acting together (e.g., although copper is an excellent electrical conductor, one copper atom alone cannot conduct electricity).

"Nano" is the SI prefix meaning "one billionth", so a nanometre is one billionth the length of a metre. The full stop at the end of this sentence is approximately 1 000 000 nanometres wide. Many materials display quite different properties when divided into grains or pieces smaller than 100 nanometres.

A sugar molecule is approximately one nanometre in width, and even some atoms have a radius of more than one quarter of a nanometre. At the scale of nanomaterials, atoms and molecules are large enough that materials cannot be treated as continuous substances. The number and arrangement of the individual atoms or molecules becomes relevant.

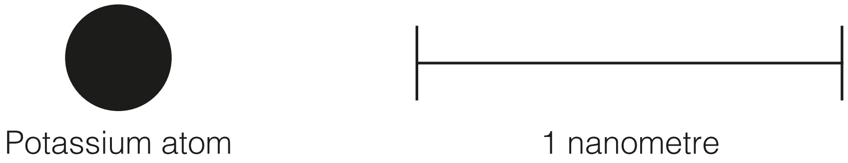

Figure 1.08: Scale of a single potassium atom

A defining feature of nanomaterials is their enormous surface area to volume ratio; greater surface area allows for more efficient catalysis of chemical reactions.

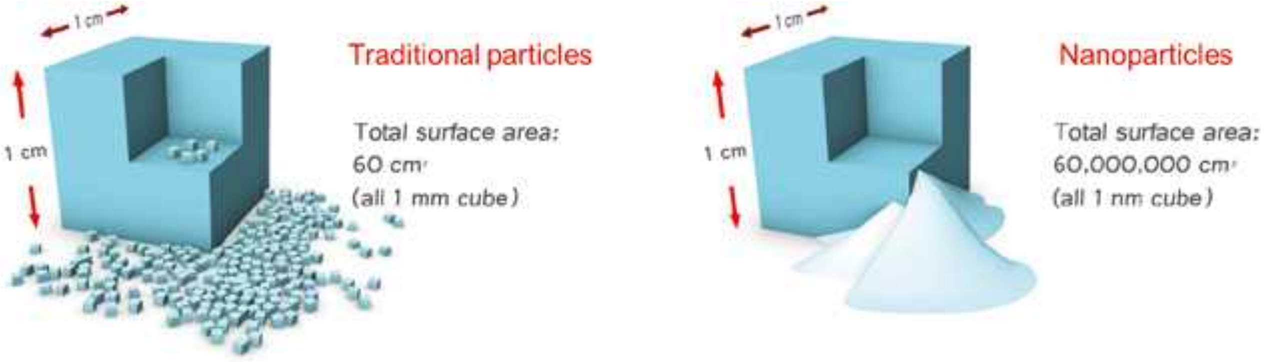

Figure 1.09: Comparison of the surface area to volume ratio of traditional materials and nanoparticles

Science as a human endeavour

Contemporary Nanotechnology Applications

Applications of nanoscale materials are many and varied.

In consumer electronics, many modern screens use Organic Light-Emitting Diode (OLED) technology, which is made from nanostructured polymer films. Unlike earlier technologies, which used colour pixels in front of a backlight, each pixel in an OLED screen emits red, green, or blue light independently. This allows for much deeper blacks: when a pixel is off, it emits no light.

In the environmental sphere, nanoscale filters are being developed which can absorb and remove large amounts of pollutants from drinking water. This reduces the amount of money and material required at large-scale water treatment plants. Large 'sponges' which can absorb many times their own weight in oil can also be used to aid clean-up efforts after major oil spills.

Some nanoparticles have useful medical properties. Silver nanoparticles can kill bacteria without harming larger organisms and have been included in some bandages. Iron oxide nanoparticles have been found to bind very well to cancer cells, which allows tiny cancerous tumours to be seen on MRI scans much more easily.

Nanomaterial science is still in its infancy, and it is usually difficult or impossible to predict the properties of a nanomaterial based on its structure. Scientists in many fields are working on measuring the properties of different nanomaterials to find useful applications.

Example

In its bulk state, gold is a yellow, non-reactive metal. However, nanoparticles of gold can be wildly different colours depending on their specific size, including red, green, and purple. Gold nanoparticles can also be used as a catalyst, and this too depends on specific size: a nanocluster made of 28 gold atoms might catalyse one reaction, while a nanocluster made of 38 gold atoms might catalyse a different reaction.

Figure 1.10: Au_{28} cluster and Au_{38} cluster

Science as a human endeavour

Virus Nanoparticles

A single virus particle is usually between 20–200 nanometres in size. Viruses straddle the border between living and non-living: although they are able to reproduce very effectively in the right conditions, they fulfill very few of the other characteristics of life. Some scientists argue that viruses should be definitively classified as non-living, which would make them natural nanomaterials—the most dangerous group of nanomaterials ever discovered!

Synthetic lipid nanoparticles were critical in the development of the mRNA vaccines which received widespread attention during the COVID-19 pandemic. The nanoparticles are used as an 'envelope' for the delivery of mRNA molecules into cells. In the absence of these envelopes, the highly unstable mRNA would break down too quickly for it to be effective at initiating an immune response.

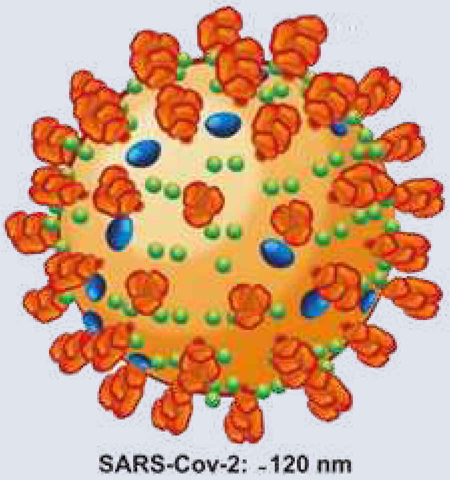

Figure 1.11: The virus that causes COVID-19 has a diameter of approximately 120 nm

Questions

4. Most commercially available sunscreens include tiny grains of titanium dioxide with diameters in the range 6–108 nm. Titanium dioxide is called a "physical blocker" because it prevents UV radiation from reaching the skin by reflecting and scattering it.

    (a) Explain why the titanium dioxide is considered to be a nanomaterial.

    (1 mark) **KA1**

    (b) Identify the property of titanium dioxide nanoparticles that makes them useful in sunscreens.

    (1 mark) **KA1**

5. The airline industry would like to consume less jet fuel to save on costs and reduce its emissions of carbon dioxide. It has been discovered that incorporating cobalt nanoparticles into the aluminium used to make aeroplanes can increase its strength by up to 100 times.

    (a) Describe the range of possible size of a cobalt nanoparticle.

    (1 mark) **KA1**

    (b) Explain how increasing the strength of the aluminium used to make the aeroplane fuselage could lead to lower jet fuel consumption in the long-term.

    (2 marks) **KA2**

6. Carbon nanotubes have excellent electrical conductivity and are extremely thin and long. They have been incorporated into some solar panels to increase the efficiency of electricity generation.

    (a) Provide a definition of the term 'nanotechnology'.

    (2 marks) **KA1**

    (b) Identify and explain one benefit of nanotechnology to consumers.

    (2 marks) **KA1**

Science understanding

Differences in the properties of substances in a mixture can be used to separate them.

- Identify how the components of a mixture can be separated by methods including filtration, distillation, and evaporation.

All materials are classified as either pure substances or mixtures.

Pure substances are composed of one type of particle. **Elements** are pure substances that contain only one type of atom (e.g., helium, iron, nitrogen), while **compounds** are pure substances containing different types of atoms that are chemically bonded in fixed ratios (e.g., water, methane, sodium chloride).

Substances that are not pure are called mixtures. **Mixtures** contain more than one element or compound that have been physically combined, but which have *not* reacted chemically with one another. Mixtures are classified as either homogenous or heterogeneous.

Homogenous mixtures are thoroughly mixed. From the Greek, *homo* means 'same', and each part of a homogenous mixture is similar in composition and properties to every other part of the mixture, even at the scale of atoms and molecules. Examples of homogenous mixtures include saltwater, alloys, and air.

Heterogeneous mixtures are incompletely mixed. Also from the Greek, *hetero* means ‘different’, and different parts of a heterogeneous mixture are distinguishable from each other in composition and character. Examples of heterogeneous mixtures include soil, salad dressing, and crude oil. Some biological mixtures can appear to be homogenous at the macro scale but are actually heterogeneous at the scale of atoms and molecules.

Example

Blood seems uniform to the naked eye, but the presence of different cells means that not all parts of the mixture are identical: red blood cells contain different elements to white blood cells. Because the components are not thoroughly mixed at the molecular level, the mixture is heterogeneous.

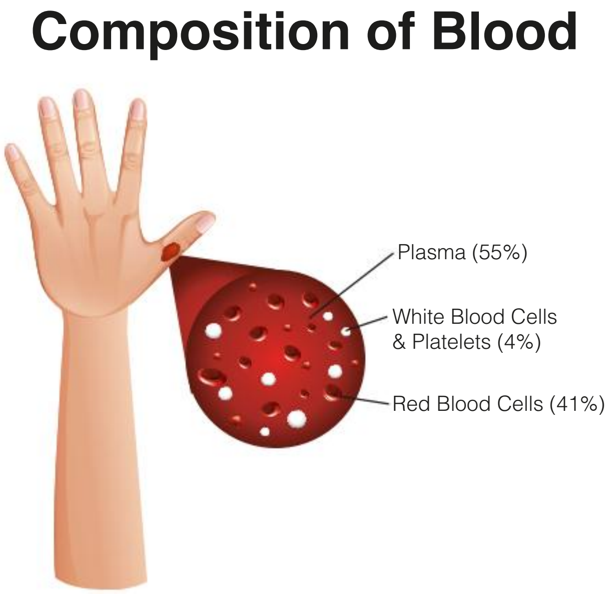

Figure 1.12: Cellular composition of blood

Mixtures, both homogenous and heterogeneous, can be separated by exploiting differences in their component substances.

Filtration

Filtration is used to separate heterogeneous mixtures that contain solid particles in a liquid medium. The mixture is passed through a **filter**, which has tiny openings that are large enough to allow liquid molecules to pass through while preventing the passage of the larger solid grains. Both the solid and the liquid components can be **isolated** and retained.

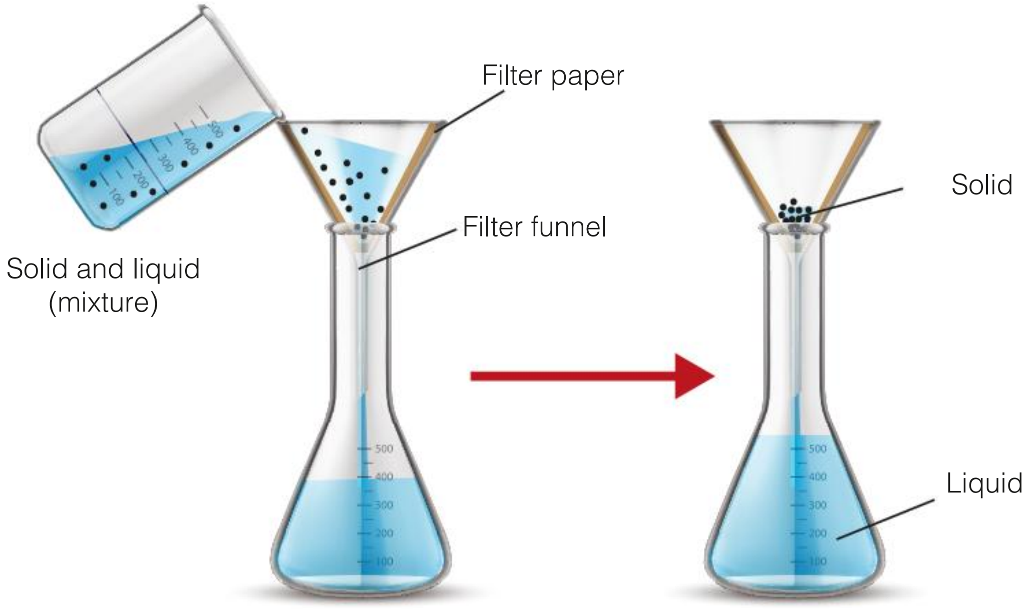

Figure 1.13: Laboratory setup for simple filtration

Evaporation

Evaporation exploits differences in boiling points to separate a liquid solvent from one or more solutes. The mixture is heated near to or above the boiling point of the solvent and the solvent evaporates. The solute, which has a much higher boiling point, remains in the evaporating dish. The solute(s) are retained but the solvent is lost.

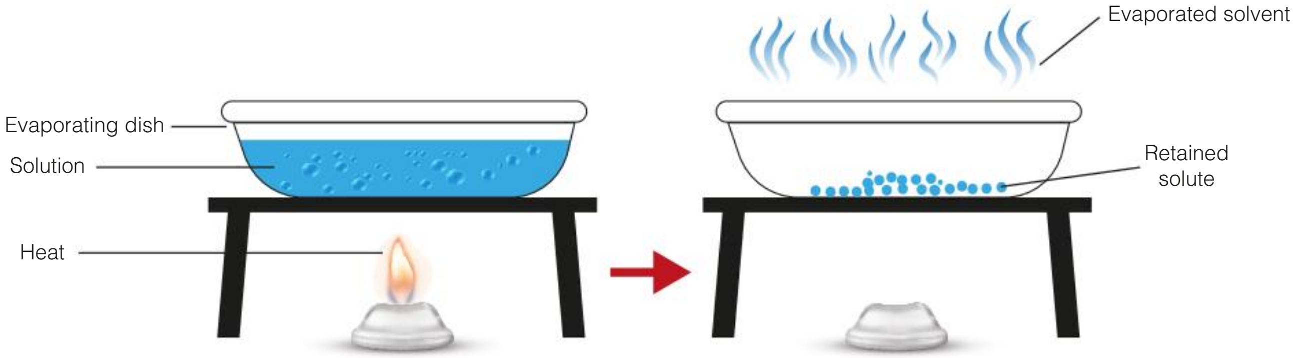

Figure 1.14: Laboratory setup for simple evaporation

Simple Distillation

Simple distillation is used to separate a liquid with a low-boiling point from other substances. Heat is transferred to the mixture and the component with the lowest boiling point vaporises. The vapour enters a condenser where it cools into a liquid and is collected; this is called the **distillate**. The low-boiling point liquid is isolated, and the other components of the mixture are retained.

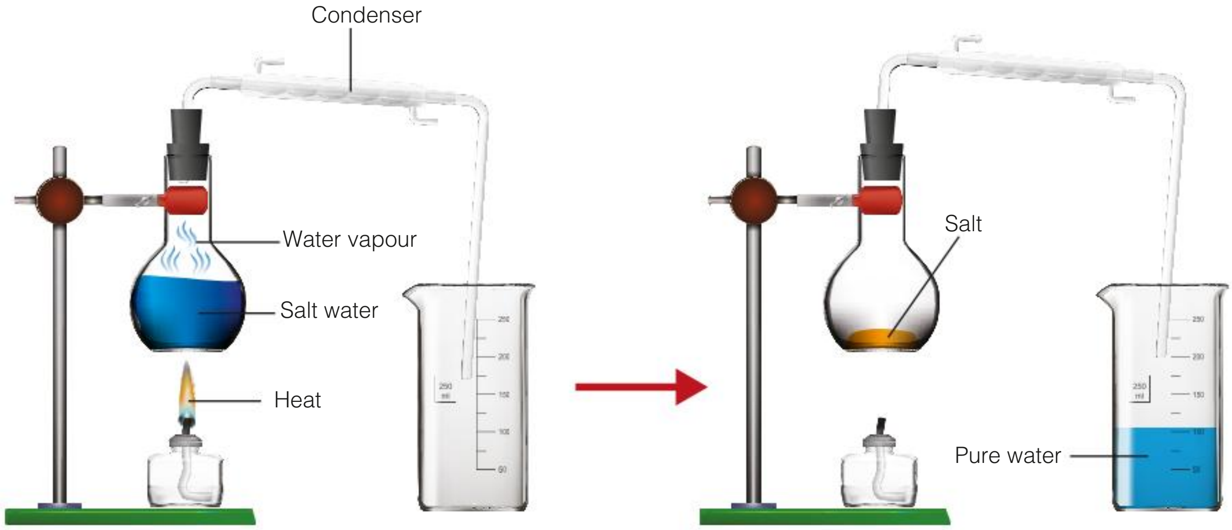

Figure 1.15: Laboratory setup for simple distillation

Science as a human endeavour

Cannabidiol Oil

Cannabidiol (CBD), a compound found in the hemp plant, is used to treat epilepsy. It can be extracted from the plant using **steam distillation**. Water is heated to generate steam, which is passed through the plant matter to vaporise cannabidiol (and some other cannabinoids). The extracts are carried with the steam into a condenser where they reform into liquids and are collected in a flask.

A similar process can be used to extract essential oils from many plants.

Figure 1.16: Steam distillation apparatus

Fractional Distillation

Fractional distillation allows liquids with similar boiling points to be separated. The technique is similar to simple distillation, but the apparatus includes a **fractionating column** with increased internal surface area for the condensation of vapours.

Figure 1.17: Fractionating columns contain beads, rings, or warped surfaces for greater surface area

The use of a fractionating column allows for much greater control over the process as the distillate flows into the condenser. The receiving flask can be swapped so as to collect different substances in different flasks. This technique is used to separate the components of crude oil.

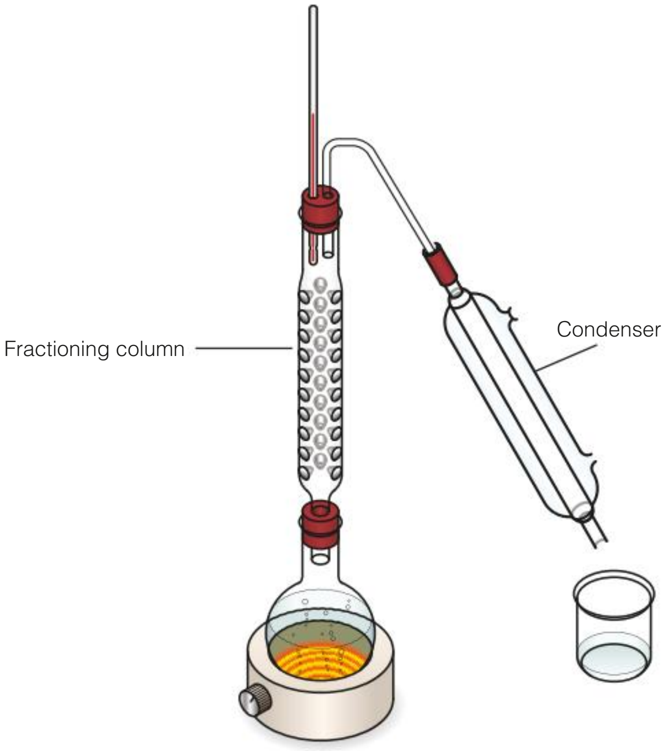

Figure 1.18: Laboratory setup for fractional distillation

Questions

7. Classify each of the following as a pure substance or mixture. For each substance, classify it as either an element or compound. For each mixture, classify it as either homogeneous or heterogeneous.

	Substances		**Mixtures**	
	Element	**Compound**	**Homogenous**	**Heterogeneous**
a lead weight	✓			
clear apple juice				
baking soda ($NaHCO_3$)				
air				
helium				
beach sand				
concrete				
carbon dioxide				
milk				
sugar ($C_{12}H_{22}O_{11}$)				
ice cream sundae				

(10 marks) **KA1**

8. Figure 1.19 shows a mixture containing potassium nitrate dissolved in water and a suspension of lead iodide, the yellow solid.

Figure 1.19: A mixture of aqueous potassium nitrate, liquid water, and solid lead iodide

(a) State and explain whether the mixture is homogenous or heterogeneous.

(2 marks) **KA1**

(b) Identify and describe a method of separating lead iodide from the mixture.

(3 marks) **KA1**

(c) The boiling points of water and potassium iodide are 100 °C and 1330 °C respectively. Describe and explain how evaporation can be used to separate the two materials.

(2 marks) **KA1**

(d) Name one other method that could be used to separate the water from the potassium iodide.

(1 mark) **KA1**

9. Methanol (wood alcohol) is a poisonous by-product of the fermentation process that produces ethanol (drinking alcohol). When spirits such as vodka are distilled, the first part of the distillate can contain a significant proportion of methanol and is therefore discarded. Some properties of methanol and ethanol are shown below.

Property	Methanol	Ethanol
Flammability	Very flammable	Very flammable
Density	0.792 g/cm^3	0.789 g/cm^3
Electrical Conductivity	Not conductive	Not conductive
Boiling Point	65 °C	78 °C
Appearance	Clear colourless liquid	Clear colourless liquid

(a) Identify the property that is used to separate methanol and ethanol during distillation.

(1 mark) **KA1**

(b) Describe the process of distillation.

(2 marks) **KA1**

(c) After discarding methanol, the ethanol distillate that is obtained is a pure substance. Provide a definition of a pure substance.

(2 marks) **KA1**

1.2 Atomic structure

Science understanding

All materials consist of atoms.

Atoms are commonly modelled as consisting of electrons orbiting a nucleus containing protons and neutrons.

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Atomic Theory

The modern understanding of the fine structure of matter has developed significantly since the atom was first proposed in Ancient Greece. The model of the atom presented by Ernest Rutherford and Niels Bohr in 1913 is the most widely referenced in chemistry classrooms today. Research chemists also use newer, more complex models which incorporate quantum theory. Nevertheless, the **Rutherford-Bohr model** is sufficient to explain most observable properties of matter, and for the purposes of this course it will be used exclusively.

Atoms are made up of **protons**, **neutrons**, and **electrons**; these are called **subatomic particles**.

Protons and neutrons exist in a central **nucleus**. The atomic nucleus is small, dense, and contains nearly all the mass of the atom. Electrons exist outside the nucleus and move very rapidly, creating an “electron cloud” around the nucleus.

The number of protons in the nucleus of an atom is called the **atomic number**; this number gives the atom its identity.

The total number of protons and neutrons in an atom is called the **mass number**. The mass number is measured in **atomic mass units** (amu), where 1 amu is equal to the mass of one proton or neutron. As electrons have approximately 1800 times less mass than protons or neutrons, they are usually ignored when determining the mass of an atom.

An atom’s **charge** is determined by the number of protons and electrons present, as neutrons do not carry electric charge. Atoms usually have an equal number of protons and electrons. Atoms therefore usually have zero electric charge overall, because protons and electrons have opposite electric charges which cancel each other.

The properties of the subatomic particles are summarised in the table below.

Particle	Symbol	Electric charge (e)	Relative mass (amu)	Position
Proton	p^+	+1	1	Nucleus
Neutron	n^0	0	1	Nucleus
Electron	e^-	–1	$\approx \frac{1}{1800}$	Around the nucleus

The electron cloud is organised into shells and each electron exists in a specific shell. These shells are often depicted in a similar way to planetary orbits, but they exist in three dimensions (hence the name 'shell').

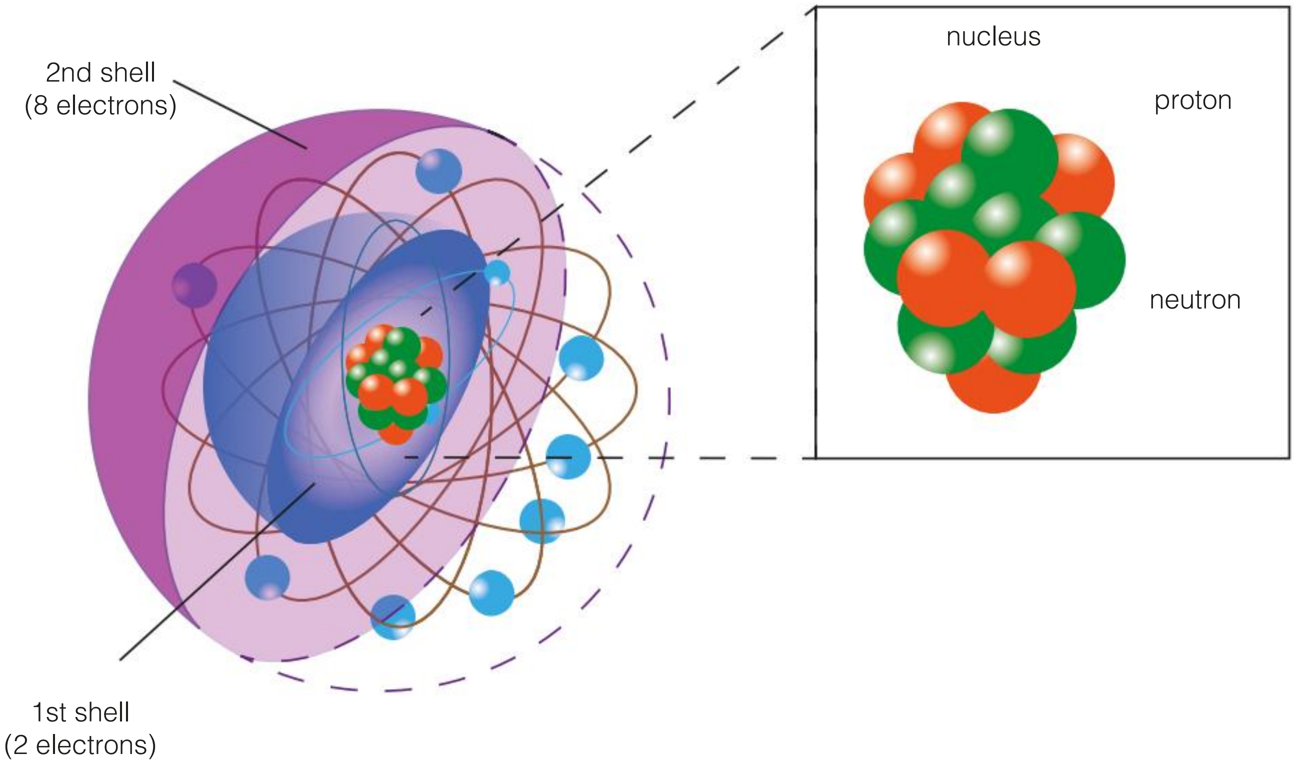

Figure 1.20: Three-dimensional diagram of an atom

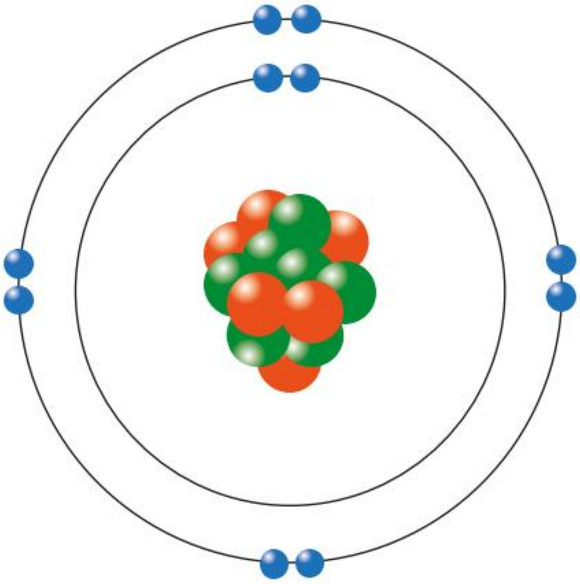

Figure 1.21: Simplified two-dimensional diagram of atom using the planetary model

Each shell has a capacity of $2n^2$ electrons where n is the number of the shell. The capacity of the first four shells is shown in the table below.

Shell	Electron Capacity
1	2
2	8
3	18
4	32

As the number of electrons in an atom increases from 1 to 18, the electrons fill the shells from the first shell outwards. The number of electrons in each shell can be described using the **main-shell electron configuration**.

Element	Number of Electrons	Main-Shell Electron Configuration
Hydrogen	1	1
Helium	2	2
Lithium	3	2,1
Beryllium	4	2,2
Boron	5	2,3
Carbon	6	2,4
Nitrogen	7	2,5
Oxygen	8	2,6
Fluorine	9	2,7
Neon	10	2,8
Sodium	11	2,8,1
Magnesium	12	2,8,2
Aluminium	13	2,8,3
Silicon	14	2,8,4
Phosphorus	15	2,8,5
Sulfur	16	2,8,6
Chlorine	17	2,8,7
Argon	18	2,8,8

Example

A sodium atom has 11 electrons. Two electrons fit into the first shell, eight more fill the second shell, and the single remaining electron exists in the third shell. The main-shell electron configuration of sodium is 2,8,1.

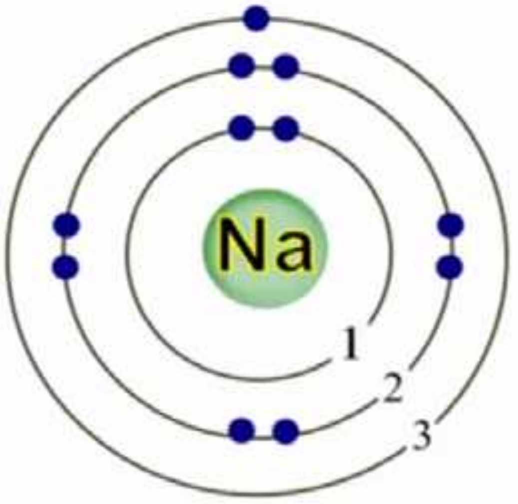

Figure 1.22: Sodium atom showing location of electrons in shells

Questions

10. Nitrogen has an atomic number of 7. It is found in the upper right section of the periodic table.

   (a) State the number of protons and electrons found in one atom of nitrogen.

(2 marks) **KA1**

   (b) Explain why the electrons in nitrogen are attracted to the nucleus.

(2 marks) **KA1**

   (c) Draw a labelled diagram of a nitrogen atom that contains 7 neutrons. Show the position of protons and neutrons and the arrangement of electrons.

(3 marks) **KA1**

   (d) Write the main-shell electron configuration of nitrogen.

(1 mark) **KA1**

Science understanding

Emissions and absorption spectra of elements provide evidence that electrons are arranged in distinct energy levels and can be used to identify some elements in matter.

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Atomic Absorption

Electrons remain associated with their nuclei because they have negative charges and are attracted to positively charged protons by electrostatic force. Just as planets orbiting a star have gravitational potential energy, electrons have **electric potential energy**. The closer an electron is to the nucleus, the less electric potential energy is stored. Each shell therefore corresponds to a specific **energy level**, as well as a distance from the nucleus. Electrons in the first shell have the lowest energy, and each electron increases in potential energy as further shells are filled.

Atoms can absorb energy, like light or heat, by promoting an electron to a higher energy level (a more distant electron shell) in a process called **atomic absorption**. The energy levels for each atom are fixed, and electrons cannot exist between the energy levels, so atoms can only absorb energy in ‘chunks’ which are equal to the amount required to promote an electron to a particular higher energy shell. These ‘chunks’ of electromagnetic energy are called **photons**.

Photons cannot be split or partially absorbed, so a photon with an amount of energy that does not match the difference in energy levels cannot interact with an atom. It is, however, possible for an electron to be promoted by more than one energy level at once.

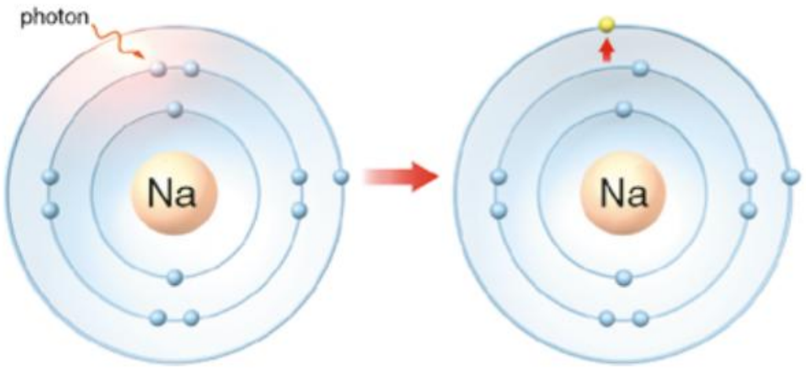

Figure 1.23: Atomic absorption of energy causes electron promotion to a higher shell

Absorption Spectroscopy

A **spectrum** is a graph, or other visual depiction, of the intensity of electromagnetic (EM) radiation at different wavelengths. The study of the spectra produced when matter interacts with or emits electromagnetic radiation is called **spectroscopy**. Visible light is one type of EM radiation that is commonly used for spectroscopic studies, but spectroscopy can also be used to study other forms of electromagnetic radiation.

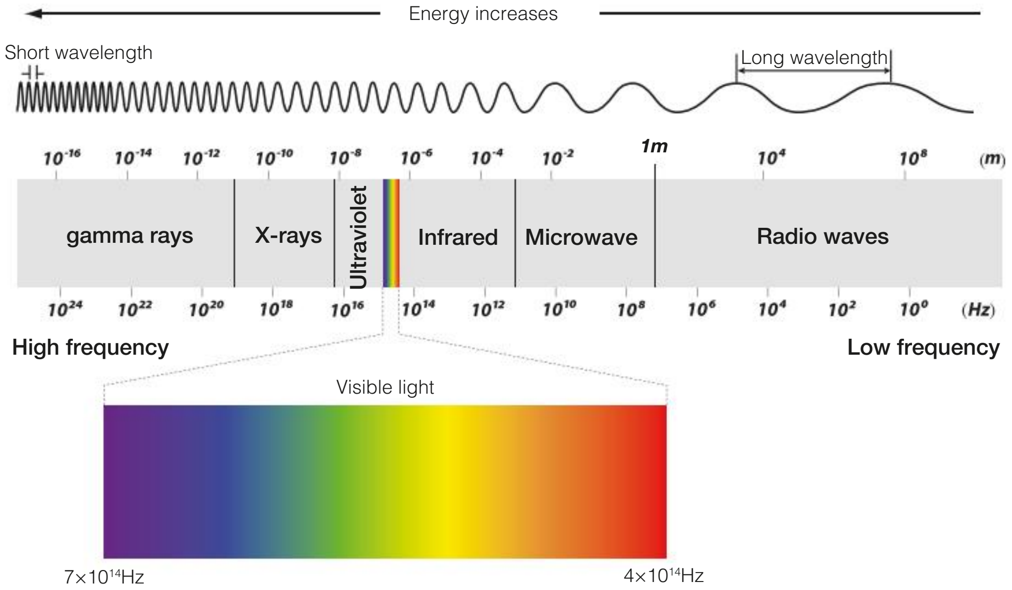

Figure 1.24: The electromagnetic spectrum

The energy carried by a photon is measured in **electron volts** (eV). Photon energy is inversely proportional to the wavelength of the photon: if the energy is doubled, the wavelength is halved. Photon wavelengths and energies also correspond to the colour of the light: low energy photons have long wavelengths and appear red, while high energy photons have short wavelengths and appear purple.

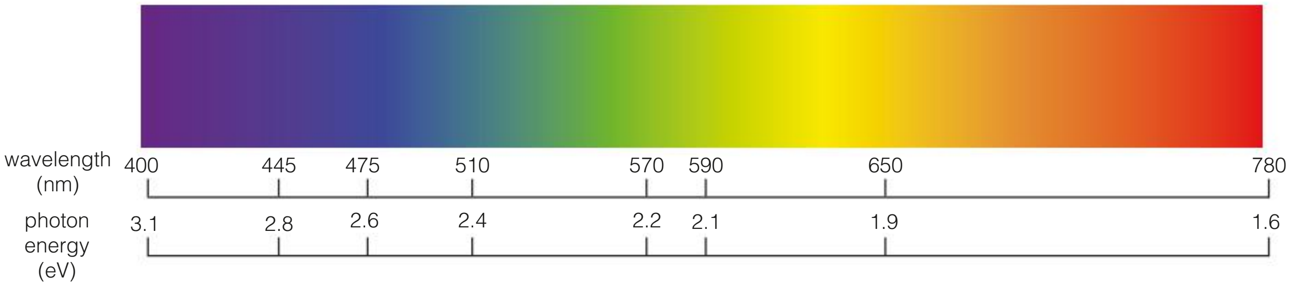

Figure 1.25: The visible portion of the EM spectrum

The specific amounts of energy that can be absorbed by an atom correspond to particular wavelengths and therefore to particular colours. Atoms will only absorb certain colours of light, while other photons with different amounts of energy will pass through. The specificity of this process is very useful for analytical purposes.

Each atom has characteristic energy levels and gaps between those energy levels. For example, all hydrogen atoms will have the same energy difference between their first and second shells, but that energy difference will not be the same as the difference between the first and second shells of a helium atom. The wavelengths of light that each atom will absorb can be used to identify the atom.

When white light consisting of photons with all visible wavelengths is shone through a sample of pure elemental gas, most of the light will pass through, but specific colours will be absorbed. If the light emerging from the sample is separated into its component wavelengths, an **atomic absorption spectrum** is obtained. The absorption spectrum will show dark lines corresponding to the differences between energy levels in that element. The position of the lines is unique and consistent for each element and can therefore be used to identify the element by comparison to known samples.

Example

Hydrogen has one electron. By absorption of a photon with 1.89 eV of energy (i.e., a photon of the correct colour), that electron can be promoted from the second shell to the third shell. The absorbed photons are no longer present in the light that is transmitted through the hydrogen gas. The absorption spectrum of hydrogen is consistent and characteristic of the element.

Hydrogen Absorption Spectrum

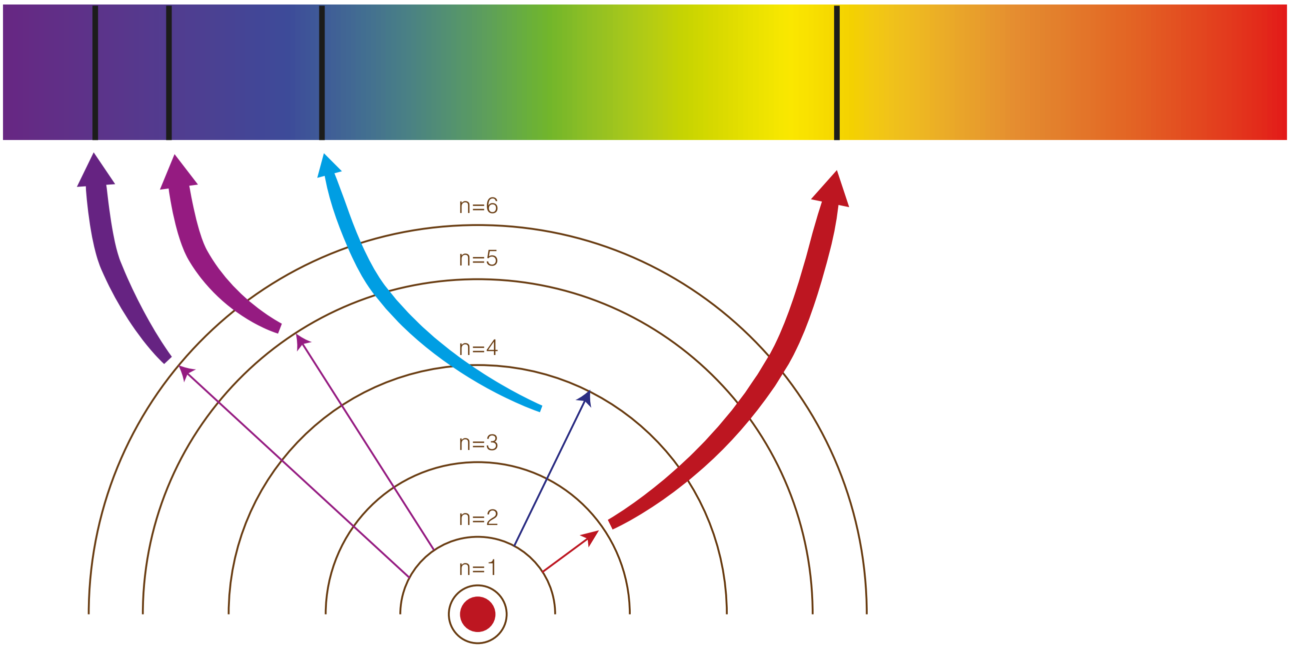

Figure 1.26: Atomic absorption wavelengths correspond to specific gaps in electron energy levels

Spectroscopy is such an exceptionally useful technique that it has been used for a broad range of applications and spawned many specialties focused on particular techniques or parts of the EM spectrum. Spectroscopic methods have been used to study the composition of distant stars, monitor dissolved oxygen content everywhere from blood to aquatic ecosystems, detect toxic compounds in food, and much else.

Science as a human endeavour

Exoplanet Spectroscopy

Exoplanets—planets which orbit stars other than the Sun—are detected by observing a very slight decrease in brightness as the exoplanet transits in front of the star. Scientists are currently attempting to measure the presence of elements and compounds in exoplanetary atmospheres using **transit spectroscopy**, which compares the relative amounts of different wavelengths coming from the star during the transit to the usual spectrum emitted by the star. Any differences detected must be due to atoms in the exoplanetary atmosphere absorbing photons with specific energies. As stars are so much larger than planets, and the amount of light absorbed so small, very sensitive instruments are required.

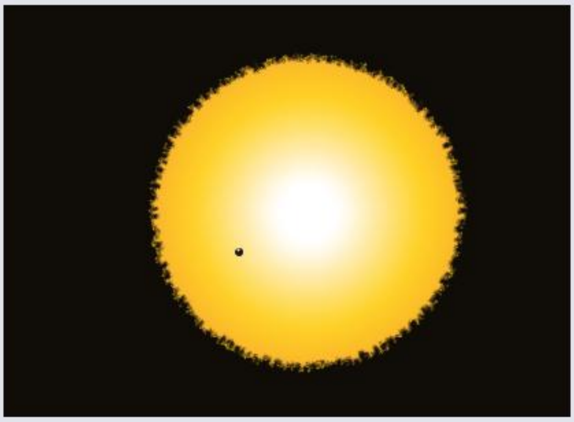

Figure 1.27: An exoplanet transiting a distant star (artist's impression)

Atomic Emission and Emission Spectroscopy

Normally, atoms exist in their **ground state**. An atom with at least one electron in a higher energy level is said to be in an **excited state**. Atoms in an excited state will have different electron configurations than the ground-state configuration.

Example

The ground-state main-shell electron configuration of sodium is 2,8,1. If the atom has absorbed a photon and an electron has been promoted, the main-shell electron configuration might instead be 2,7,2. Electron configurations are explained in detail later in this chapter.

Atoms in an excited state are unstable, and the electrons return to their original energy levels within a fraction of a second. As they do so, they emit a photon with energy equal to the energy that was originally absorbed.

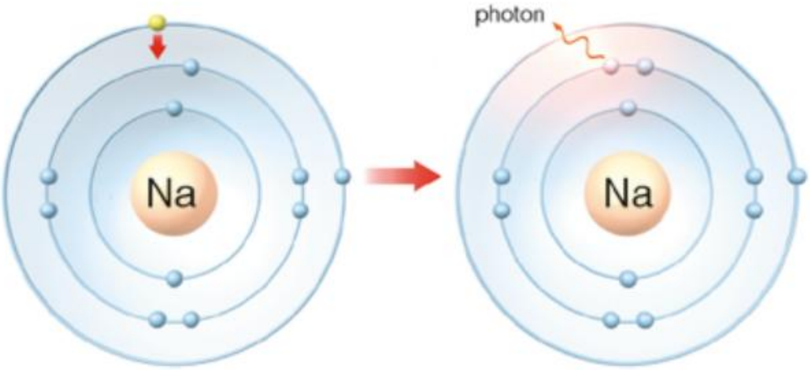

Figure 1.28: Atomic emission of energy as an electron returns to its ground state

Separating the emitted light into its component wavelengths produces an **emission spectrum**. Because the differences between energy levels in an atom are fixed, the wavelengths in the emission spectrum of a given element are identical to the wavelengths of its absorption spectrum.

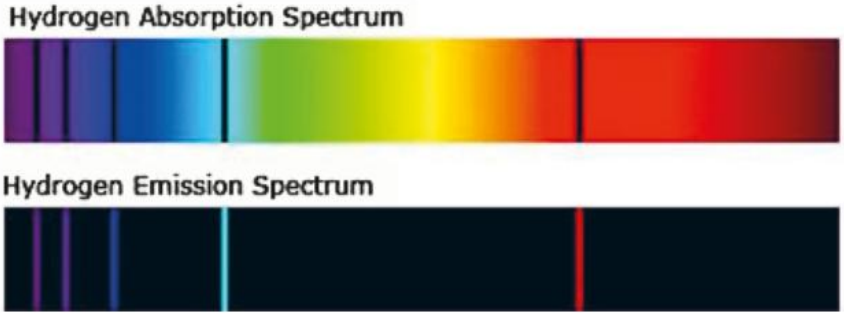

Figure 1.29: Absorption and emission spectra of hydrogen

When more than one electron is included in the system, the number of possible jumps between different energy levels increases rapidly. Most elements have complex emission spectra.

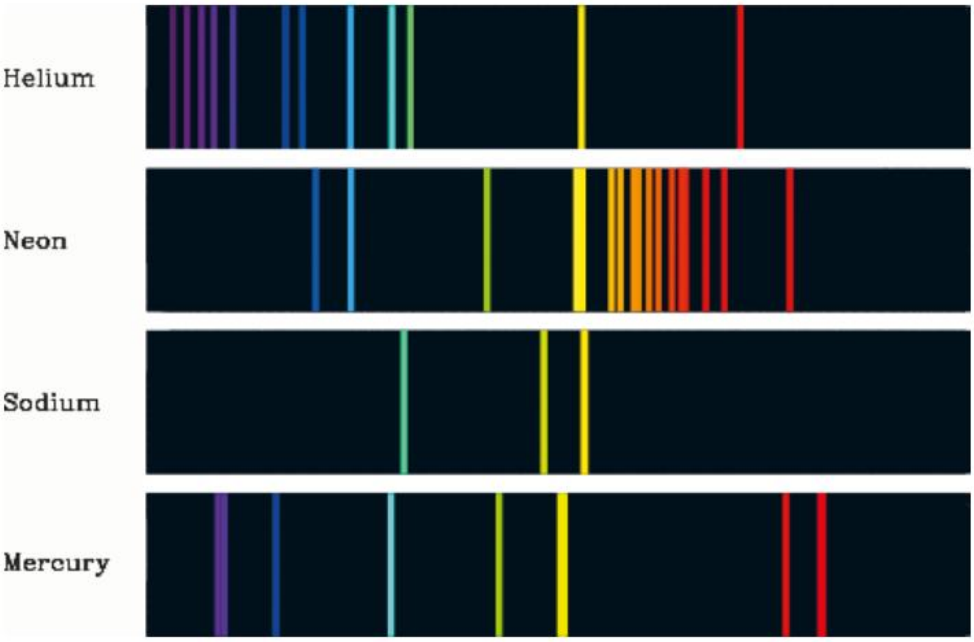

Figure 1.30: Emission spectra for helium, neon, sodium, and mercury

Questions

11. Describe how atoms can absorb certain quantities of energy but not others.

....

....

.... *(2 marks)* **KA1**

12. The absorption and emission spectrum of each element is characteristic for that element. Explain what is meant by 'characteristic' in this context.

....

....

.... *(1 mark)* **KA4**

13. In 1815, Joseph von Fraunhofer discovered dark lines in the spectrum of visible light coming from the Sun.

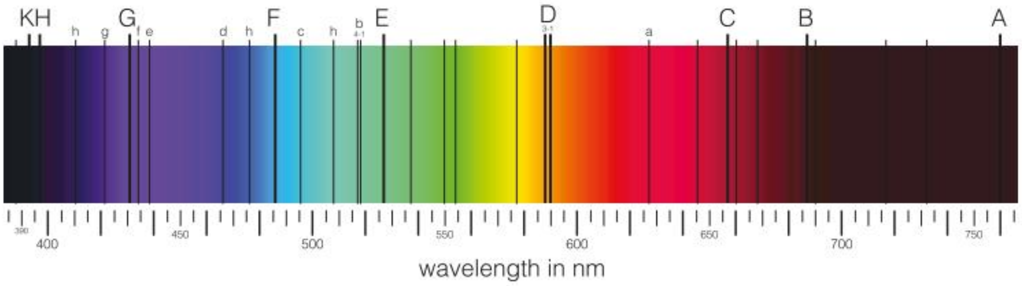

Figure 1.31: Solar spectrum

(a) Name the process that causes the dark lines in the spectrum of visible light emitted by the Sun.

.... *(1 mark)* **KA1**

(b) Explain how the dark lines provide evidence that electrons in atoms are arranged in distinct energy levels.

....

....

....

.... *(3 marks)* **KA1**

(c) Describe a method of using the spectrum above to determine whether magnesium and iron are present in the atmosphere of the Sun.

(3 marks) **KA1**

Science understanding

Atomic number and mass number provide information about the numbers of subatomic particles in an atom.

Many elements consist of a number of different isotopes, which have different physical properties but the same chemical properties.

- Represent isotopes of an element using appropriate notation.

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An individual atom can be represented using a composite symbol with three parts. The elemental symbol (X) allows identification of the element without reference to the periodic table. The atomic number (Z) is the number of protons present in the nucleus. The mass number (A) is the sum of protons and neutrons in the nucleus.

$${}^{A}_{Z}\mathrm{X}$$

The number of neutrons present in an atom can be calculated by subtracting the atomic number from the mass number.

Example

$${}^{42}_{20}\mathrm{Ca}$$

Like all calcium atoms, the atom of calcium represented above has 20 protons. It has a mass number of 42, so there must also be 22 neutrons present in the nucleus.

$$\begin{aligned}\text{number of neutrons} &= A - Z\\ &= 42 - 20\\ &= 22\end{aligned}$$

Isotopes

An **isotope** of an element is a configuration of a nucleus with a defined number of neutrons, and therefore with a defined mass number. The number of protons and electrons does not change between isotopes of an element. Isotopes are identified using the elemental name and the mass number. The various isotopes of an element have identical chemical properties because these are determined by the behaviour of the electrons.

Example

Nitrogen has two stable isotopes. Nitrogen-14 has seven protons and *seven* neutrons, while nitrogen-15 has seven protons and *eight* neutrons.

$${}^{14}_{7}\mathrm{N} \qquad {}^{15}_{7}\mathrm{N}$$

An atom with eight protons would be an isotope of oxygen.

Most elements have at least two isotopes, some of which are unstable. Unstable isotopes **decay** into stable isotopes by emitting energy or particles from their nucleus. Nitrogen has 14 known unstable isotopes, with mass numbers in the range 10–25.

Eighty elements have stable isotopes; these are elements 1–42, 44–60, and 62–82. Element 43 (technetium), element 61 (promethium), and all elements with atomic number greater than 82 have no stable isotopes.

Example

Of the oxygen atoms on Earth, 0.2% are oxygen-18.

$${}^{18}_{8}\mathrm{O}$$

number of neutrons = A – Z
= 18 – 8
= 10

There are 10 neutrons in an atom of oxygen-18.

Questions

14. Use the periodic table to complete the following table.

Name	Symbol	Atomic Number (Z)	Mass Number (A)	Number of protons	Number of electrons	Number of neutrons
Sodium	${}^{24}_{11}\mathrm{Na}$	11	24			
	${}^{89}_{\_\_}\mathrm{Y}$			39	39	
Copper						35
	${}^{207}_{\_\_}\mathrm{Pb}$					
Thallium		81	204			
	${}^{1}_{1}\mathrm{H}$					
Carbon			12			
	${}^{\_}_{\_}\mathrm{N}$					7
Barium			137			
Calcium						20

(10 marks) **KA1**

15. State the number of protons, electrons, and neutrons in the following elements.

Argon-36 ....

Boron-11 ....

Neon-22 ....

(3 marks) **KA1**

16. Technetium (Tc) has an atomic number of 43. Technetium-99 is an isotope of technetium that is widely used in nuclear medicine.

    (a) Calculate the number of neutrons in technetium-99.

    (1 mark) **KA1**

    (b) Represent technetium-99 using a chemical symbol.

    (1 mark) **KA1**

17. Element 78, platinum, has six naturally occurring isotopes.

    (a) Define the term isotope.

    (1 mark) **KA1**

    (b) State why the isotopes of an element have identical chemical properties.

    (2 marks) **KA1**

    (c) Complete the table below.

Isotope	Number of protons	Number of electrons	Number of neutrons
platinum-190			
platinum-192			
platinum-194			
platinum-195			
platinum-196			
platinum-198			

(6 marks) **KA1**

Science understanding

The arrangement of electrons in atoms and monatomic ions can be described in terms of shells and subshells.

- Write the electron configuration using subshell notation of an atom of any of the first 38 elements in the periodic table.

Electron shells have so far been presented as circular and uniform, but this is a simplification. Each shell is made up of **subshells**, and the number of subshells in a shell is equal to the shell number (that is, the first shell has one subshell, the second shell has two subshells, etc).

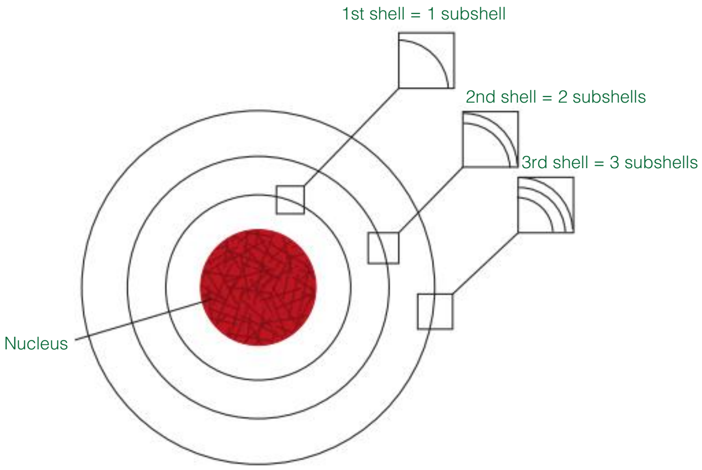

Figure 1.32: Each additional shell contains one additional subshell

Subshells are named using the number of the main shell and the letters *s*, *p*, *d*, and *f*. The first main shell has only the 1*s* subshell, the second main shell comprises the 2*s* subshell and the 2*p* subshell, the third main shell comprises the 3*s*, 3*p*, and 3*d* subshells, and so on. The number of electrons that each subshell can hold is shown in the table below.

Subshell	Electron Capacity
s	2
p	6
d	10
f	14

Electron configuration in subshells can be represented using **subshell notation**. When the phrase 'electron configuration' is not preceded by the words 'main-shell', it always uses the system of subshell notation. Electron configurations are written using the name of the subshell and a superscript to the right of the name showing the number of electrons present in that subshell.

Example

Lithium has three electrons. Two exist in the 1*s* subshell, represented by "$1s^2$". The third electron is in the 2*s* subshell, represented by "$2s^1$". The electron configuration of lithium is:

$$1s^2 2s^1$$

Subshells within a main shell do not have equal energy levels and, beginning with the third shell, the subshells of different main shells begin to overlap. As shown in Figure 1.33, the 4*s* subshell is of lower energy than the 3*d* subshell. Shells always fill from lowest energy first, so the 19th electron in an atom will exist in the 4*s* subshell rather than the 3*d* subshell.

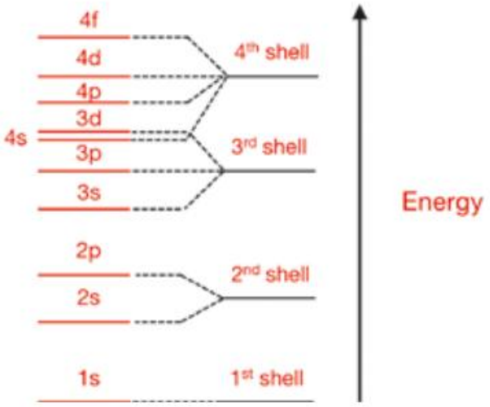

Figure 1.33: Comparative energies of the first four shells

Example

The electron configuration of element 19, potassium, is $1s^22s^22p^63s^23p^6\mathbf{4s^1}$.

This means potassium has a main-shell electron configuration of 2,8,8,1.

If subshells were ignored, a main-shell configuration of 2,8,9 might be incorrectly predicted.

The electron configurations of the first 20 elements are shown in the table below.

Element	**Z**	**Electron configuration**
Hydrogen	1	$1s^1$
Helium	2	$1s^2$
Lithium	3	$1s^22s^1$
Beryllium	4	$1s^22s^2$
Boron	5	$1s^22s^22p^1$
Carbon	6	$1s^22s^22p^2$
Nitrogen	7	$1s^22s^22p^3$
Oxygen	8	$1s^22s^22p^4$
Fluorine	9	$1s^22s^22p^5$
Neon	10	$1s^22s^22p^6$
Sodium	11	$1s^22s^22p^63s^1$
Magnesium	12	$1s^22s^22p^63s^2$
Aluminium	13	$1s^22s^22p^63s^23p^1$
Silicon	14	$1s^22s^22p^63s^23p^2$
Phosphorus	15	$1s^22s^22p^63s^23p^3$
Sulfur	16	$1s^22s^22p^63s^23p^4$
Chlorine	17	$1s^22s^22p^63s^23p^5$
Argon	18	$1s^22s^22p^63s^23p^6$
Potassium	19	$1s^22s^22p^63s^23p^64s^1$
Calcium	20	$1s^22s^22p^63s^23p^64s^2$

The third, fourth, and fifth main shells overlap significantly, so the order in which subshells are filled as further electrons are added continues to be complex.

1s 2s 2p 3s 3p 4s 3d 4p 5s

There are two straightforward methods for remembering the sequence in which subshells are filled.

Method 1: Subshells Columns with Diagonal Lines

If the subshells are written in columns corresponding to their *s*, *p*, *d*, or *f* character, descending diagonal lines show the correct order for filling.

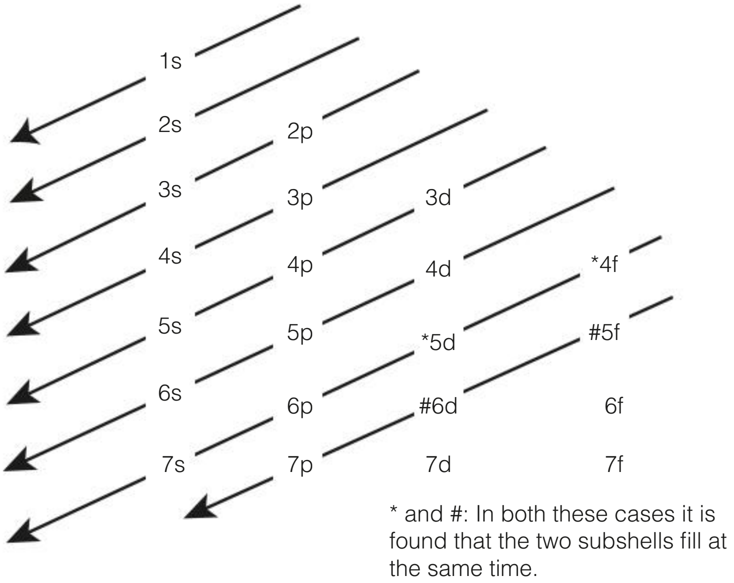

Figure 1.34: Diagonal method for determining fill order for electron subshells

Method 2: Reading Subshells from the Periodic Table

The periodic table is arranged such that the elements with *s* electrons in their highest energy subshell are grouped together in the *s*-block, elements with *p* electrons in their highest energy subshell are grouped in the *p*-block, and so on. The order of subshells can be read left-to-right, top-to-bottom, directly from the table.

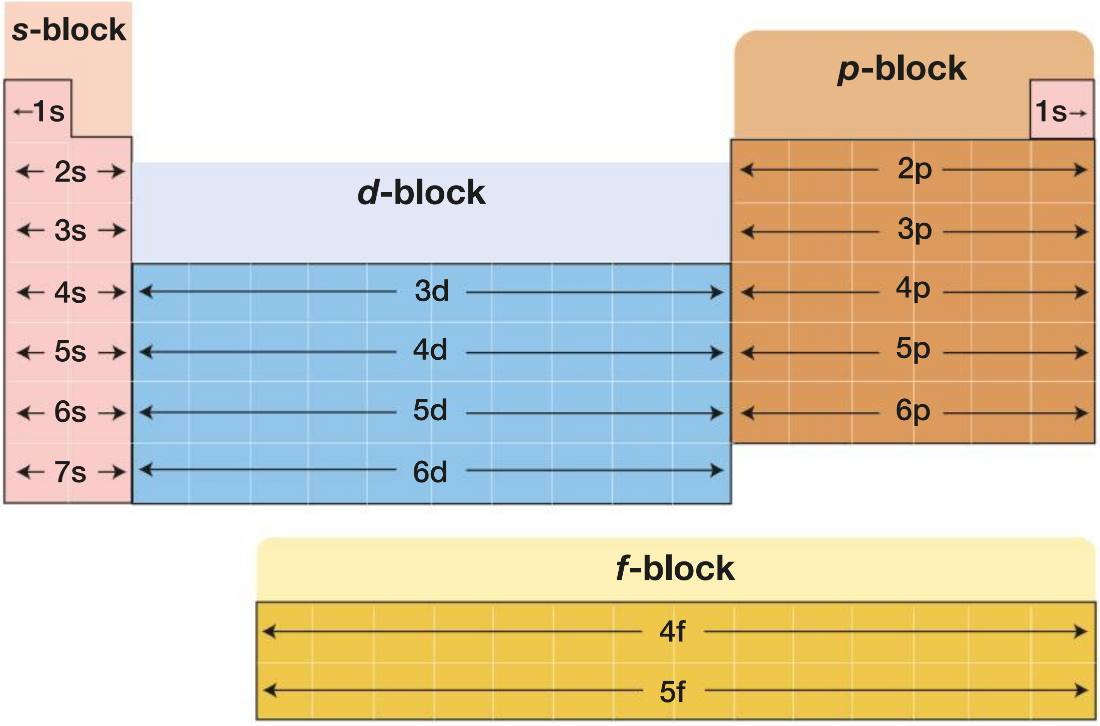

Figure 1.35: Periodic table method for determining fill order for electron subshells

Example

Arsenic is element 33. It is found in the *p*-block, and it is the third element in its row in the *p*-block, so its configuration is expected to end with 'p^3'. Using either method above for determining the order in which subshells are filled, the electron configuration for arsenic is $1s^22s^22p^63s^23p^64s^23d^{10}4p^3$.

The electron configurations of elements 21-38 are shown in the table below.

Element	Symbol	Z	Electron configuration
Scandium	Sc	21	$1s^22s^22p^63s^23p^64s^23d^1$
Titanium	Ti	22	$1s^22s^22p^63s^23p^64s^23d^2$
Vanadium	V	23	$1s^22s^22p^63s^23p^64s^23d^3$
Chromium*	Cr	24	$1s^22s^22p^63s^23p^6\mathbf{4s^13d^5}$
Manganese	Mn	25	$1s^22s^22p^63s^23p^64s^23d^5$
Iron	Fe	26	$1s^22s^22p^63s^23p^64s^23d^6$
Cobalt	Co	27	$1s^22s^22p^63s^23p^64s^23d^7$
Nickel	Ni	28	$1s^22s^22p^63s^23p^64s^23d^8$
Copper*	Cu	29	$1s^22s^22p^63s^23p^6\mathbf{4s^13d^{10}}$
Zinc	Zn	30	$1s^22s^22p^63s^23p^64s^23d^{10}$
Gallium	Ga	31	$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$
Germanium	Ge	32	$1s^22s^22p^63s^23p^64s^23d^{10}4p^2$
Arsenic	As	33	$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$
Selenium	Se	34	$1s^22s^22p^63s^23p^64s^23d^{10}4p^4$
Bromine	Br	35	$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$
Krypton	Kr	36	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$
Rubidium	Rb	37	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$
Strontium	Sr	38	$s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$

*In chromium atoms and copper atoms, one electron moves from the 4*s* subshell to the 3*d* shell, giving electron configurations ending in $4s^13d^5$ and $4s^13d^{10}$, respectively. This occurs because the presence of five or 10 electrons in the *d* subshell creates a symmetry which provides an energy benefit greater than the energy required to promote the electron by one shell. The electron configurations shown are therefore more energetically stable than the alternative $4s^2$ configurations.

Electron Configurations of Ions

Ions are charged atoms formed by the gain or loss of electrons. Metals tend to release electrons and form positive **cations**, while non-metals tend to gain electrons to form negative **anions**. Electrons are mostly gained or lost in the same order as the shells are filled.

Example

A chlorine atom (Cl) can gain one electron to become a chloride anion (Cl^-). The additional electron is added to the lowest energy subshell that is available, which is 3*p*.

Chlorine Atom (Cl)	Chloride Ion (Cl^-)
$1s^22s^22p^63s^23p^5$	$1s^22s^22p^63s^23p^6$

Example

A beryllium atom (Be) can lose two electrons to become a beryllium cation (Be^{2+}). The two electrons are lost from the highest energy subshell that already contains electrons, which is 2*s*.

Beryllium Atom (Be)	**Beryllium Ion (Be^{2+})**
$1s^2 2s^2$	$1s^2$

Ions of Transition Metals

There is one significant exception to the method of determining the electronic configurations of ions. Metals with their highest energy electrons in the *d*-block (transition metals) prefer to lose electrons from the n*s* subshell before the (n–1)*d* subshell.

Example

Nickel (Ni) can lose two electrons to become a nickel cation (Ni^{2+}). The two electrons are removed from the 4*s* subshell, rather than the 3*d* subshell.

Nickel Atom (Ni)	**Nickel Ion (Ni^{2+})**
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^8$

Questions

18. Using appropriate chemical terminology, explain the distinction between a **main shell** and a **subshell**.

(2 marks) **KA1**

19. Explain why the second period on the periodic table does not include any *d*-block elements.

(2 marks) **KA1**

20. Write the electron configurations for the following elements using subshell notation.

a. hydrogen
b. beryllium
c. carbon
d. oxygen
e. fluorine
f. sodium
g. sulfur
h. argon
i. calcium
j. chromium
k. copper
l. arsenic
m. selenium
n. chloride ion, Cl^-
o. magnesium ion, Mg^{2+}
p. phosphorus ion, P^{3-}
q. vanadium(III) ion, V^{3+}
r. cobalt(II) ion, Co^{2+}
s. zinc(II) ion, Zn^{2+}
t. oxygen ion, O^{2-}

(20 marks) **KA1**

1.3 Quantities of atoms

Science understanding

The quantities of different substances can be conveniently compared using the mole unit.

The relative atomic mass of an element is determined from all the isotopes of that element.

The number of moles of atoms in a sample can be determined from the number of atoms present or from the mass of the atoms.

- Undertake calculations using the relationship n = m/M and its rearrangements.

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Chemical reactions involve whole numbers of atoms or molecules. However, atoms are far too small to be individually counted. The primary way of quantifying a real sample of a substance in a laboratory is to measure its mass, not count the number of particles it contains. To understand and control the quantities of substances consumed and produced in a reaction, a convenient method of converting between mass and number of particles is required. The mole unit was created for this purpose.

A **mole** of a substance is the amount of that substance that contains $6.02214076 \times 10^{23}$ particles of that substance (usually rounded to 6.02×10^{23}). This is very close to the exact number of carbon atoms in a 12.00 g sample of carbon-12. It is called **Avogadro's number** (N_A) in honour of Amadeo Avogadro.

Example

Just as one might count eggs in dozens, particles can be counted in moles.

3 dozen eggs = 36 eggs

1 mole of particles = approximately 602,000,000,000,000,000,000,000 particles.

Chemists chose the number of particles in a mole to make their calculations especially easy. The mass of individual molecules is expressed in **atomic mass units** (amu), where one atomic mass unit is equal to $\frac{1}{12}$ the mass of one atom of carbon-12. Because Avogadro's number is also based on carbon-12, the mass of one atom or molecule in atomic mass units is equal to the mass of one mole of that substance in grams.

Molar Mass

The **molar mass** (M) of a substance is the mass of one mole of that substance in units of grams per mole (g/mol). The molar mass of one atom of each element is written beneath the symbol on the periodic table. Molar masses can be used to determine the amounts of reactants that should be used in a reaction.

Example

In the reaction of carbon and oxygen to form carbon dioxide, exactly one carbon atom reacts with exactly one oxygen molecule (containing two oxygen atoms) to produce exactly one carbon dioxide molecule.

$$C + O_2 \rightarrow CO_2$$

A chemist wishes to conduct this reaction in the laboratory and produce exactly one mole of carbon dioxide.

The molar mass of carbon is 12.01 g/mol.

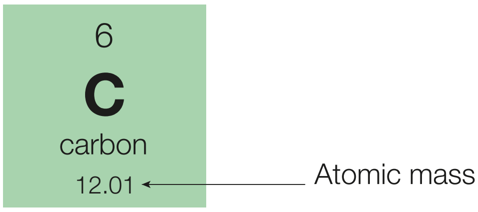


Figure 1.36: Atomic mass for each element is shown on the periodic table

One mole of carbon will weigh 12.01 grams.

$$M_C = 12.01 \text{ g/mol}$$

The molar mass of oxygen atoms is 16.00 g/mol.

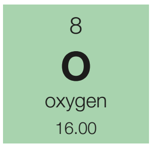


Figure 1.37: Atomic mass for each element is shown on the periodic table

An oxygen molecule contains two oxygen atoms, so the molar mass of the molecule is 32.00 g/mol. One mole of oxygen molecules will weigh 32.00 grams.

$$\begin{aligned} M_{O_2} &= (2 \times M_O) \\ &= (2 \times 16.00) \\ &= 32.00 \text{ g/mol} \end{aligned}$$

The ratio of carbon to oxygen in the reaction is 1:1, so one mole of carbon will react with one mole of oxygen to produce carbon dioxide with no reactants left over. The chemist in the laboratory knows to weigh out 12.01 grams of carbon and 32.00 grams of oxygen.

One mole of carbon dioxide will be produced. The molar mass of carbon is calculated by adding the molar masses of its constituent atoms: one carbon and two oxygen atoms.

$$\begin{aligned} M_{CO_2} &= (1 \times M_C) + (2 \times M_O) \\ &= (1 \times 12.01) + (2 \times 16.00) \\ &= 44.01 \text{ g/mol} \end{aligned}$$

The amount of carbon dioxide produced will be 44.01 g.

Questions

21. Use the periodic table to calculate the molar mass of each of the following substances.

(a) hydrogen gas, H_2

(b) ozone, O_3

(c) neon gas, Ne

(d) caffeine, $C_8H_{10}N_4O_2$

(e) lithium hydride, LiH

(f) sodium nitride, NaN_3

(g) aluminium sulfate, $Al_2(SO_4)_3$

(h) isopropyl alcohol, $CH_3CH_2CH_2OH$

(8 marks) **KA1**

Calculating with Moles

Reactions will not always occur in amounts corresponding to exactly one mole. The number of moles in a sample can be calculated using the formula below.

$$n = \frac{m}{M}$$

The number of moles present, n, is measured in moles (mol).

The mass of the sample, m, is measured in grams (g).

The molar mass of the substance, M, is measured in grams per mole (g/mol).

Example

Calculate the number of moles of iron present in an iron nail with a mass of 0.214 g.

$$\begin{aligned} m &= 0.214\text{ g} \\ M_{Fe} &= 55.85\text{ g/mol} \\ n &= \frac{m}{M} \\ &= \frac{0.214\text{ g}}{55.85\text{ g/mol}} \\ &= 0.00383\text{ mol} \\ &= 3.83 \times 10^{-3}\text{ mol} \end{aligned}$$

Example

Calculate the number of moles of methane (CH_4) present in 1.000 kg of methane.

$$\begin{aligned} m &= 1000\text{ g} \\ M_{CH_4} &= (1 \times 12.01) + (4 \times 1.01) \\ &= 16.05\text{ g/mol} \\ n &= \frac{m}{M} \\ &= \frac{1000\text{ g}}{16.043\text{ g/mol}} \\ &= 62.33\text{ mol} \end{aligned}$$

The number of moles of a particular atom in a sample will be the number of moles of the substance multiplied by the number of atoms in one molecule or formula unit.

Example

Calculate the number of moles of *hydrogen atoms* (H) present in 1.000 kg of methane (CH_4).

$$\begin{aligned} n_{CH_4} &= 62.33\text{ mol} \qquad \textit{(from above)} \\ n_H &= n_{CH_4} \times \text{number of hydrogens in } CH_4 \\ &= 62.33 \times 4 \\ &= 249.3\text{ mol} \end{aligned}$$

Moles are SI units and can take all the usual SI prefixes. Small fractions of a mole can therefore be expressed using **millimoles** (mmol), **micromoles** (μmol), or even **nanomoles** (nmol).

Example

Paclitaxel is a chemotherapy drug with a molecular formula of $C_{47}H_{51}NO_{14}$ and a molar mass of 853.91 g/mol. Doses are calculated for specific individuals, but a typical dose might be 10 milligrams (mg).

$$\begin{aligned}
m &= 10\text{ mg}\\
&= 0.01000\text{ g}\\
M &= 853.91\text{ g/mol}\\
n &= \frac{m}{M}\\
&= \frac{0.01000\text{ g}}{853.91\text{ g/mol}}\\
&= 0.00001171\text{ mol}\\
&= 11.71\ \mu\text{mol}
\end{aligned}$$

A typical dose of paclitaxel is approximately 12 micromoles.

Mass Number, Atomic Mass, and Relative Atomic Mass

The **mass number** of an atom is simply the sum of protons and neutrons. It is always a whole number.

The **atomic mass** of an atom is its precise mass measured in atomic mass units (amu). Carbon-12 has an atomic mass of exactly 12, but most other elements don't have whole-number atomic masses because of the tiny discrepancies introduced by the mass of electrons and relativistic effects.

The molar masses provided for each element on the periodic table are **relative atomic masses** (A_r). Relative atomic masses are averages calculated using the atomic masses of each isotope and weighted according to the abundance of different isotopes on Earth. Beyond Earth, there may be quite different amounts of each isotope, and the molar mass of each element would therefore also be different.

Example

On Earth, bromine occurs as two stable isotopes. Bromine-79 makes up 50.7% of bromine atoms, and bromine-81 makes up the remaining 49.3%.

The atomic mass (A) of bromine-79 is 78.92 amu.

The atomic mass of bromine-81 is 80.92 amu.

The relative atomic mass is calculated as the weighted average of these two atomic masses.

$$\begin{aligned}
A_r &= (\%\ {}^{79}\text{Br} \times A_{^{79}\text{Br}}) + (\%\ {}^{81}\text{Br} \times A_{^{81}\text{Br}})\\
&= (0.507 \times 78.92) + (0.493 \times 80.92)\\
&\approx 79.91
\end{aligned}$$

The **relative molecular mass** of a molecular compound is calculated by summing the relative atomic masses of the constituent atoms in the molecule. The **relative formula mass** of a continuous compound is also calculated by summing the relative atomic masses of the constituent atoms in the formula unit.

Relative molecular mass (M_r) and relative formula mass (M_r) are calculated in the same way as molar mass (M), and the terms are often used interchangeably. However, molar mass is the mass of one mole, and has units of grams per mole. Relative masses are compared to the mass of carbon-12 and therefore do not have units. Quantities without units are called **dimensionless** quantities.

The molar mass and relative molecular mass of water are compared below.

Molar Mass	Relative Molecular Mass
$M_{H_2O} = (2 \times 1.01\text{ g/mol}) + (1 \times 16.00\text{ g/mol})$	$M_{r(H_2O)} = (2 \times 1.01) + (1 \times 16.00)$
$= 18.02\text{ g/mol}$	$= 18.02$

Science as a human endeavour

Mass Spectrometry

Values in this text have been rounded to two or three decimal places, but the relative atomic mass of individual isotopes is known very precisely. This means that, just as atoms have characteristic absorption and emission spectra, individual molecules have characteristic masses that can be used to identify them. For example, the mass of one molecule of water containing no rare isotopes (i.e., two atoms of ^{1}H and one atom of ^{16}O) is exactly 18.010 564 683 4. This is the basis of mass spectrometry, which can rapidly detect specific substances.

Swab tests for explosives conducted at airports rely on this technique. A security officer will rub a small swab over the clothes and bags of an individual, then place the swab into a mass spectrometer that fits on a table. The surface of the swab is vaporised and the machine scans for compounds that are present in explosives. Mass spectrometers of this type usually return a result within seconds.

Questions

22. Calculate the relative molecular mass of the following compounds.

(a) ammonia, NH_3

(d) butane, C_4H_{10}

(b) nitrogen dioxide, NO_3

(e) carbon tetrafluoride, CF_4

(c) hypochlorous acid, HOCl

(f) acetone, CH_3COCH_3

(6 marks) **KA1**

23. Calculate the relative formula mass of the following compounds.

(a) calcium chloride, $CaCl_2$

(d) lithium nitride, Li_3N

(b) magnesium phosphate, $Mg_3(PO_4)_2$

(e) strontium hydroxide, $Sr(OH)_2$

(c) potassium oxide, K_2O

(f) hydrated barium chloride, $BaCl_2.2H_2O$

(6 marks) **KA1**

24. Common sugar, sucrose, has the molecular formula $C_{12}H_{22}O_{11}$.
    (a) Calculate the relative molecular mass of one sugar molecule.

(1 mark) **KA1**

    (b) Calculate the mass of 1.00 moles of sugar molecules.

(2 marks) **KA1**

    (c) State the number of sugar molecules in one mole.

(1 mark) **KA1**

    (d) State the number of moles of hydrogen atoms in one mole of sugar.

(1 mark) **KA1**

    (e) Calculate the number of hydrogen atoms in one mole of sugar.

(1 mark) **KA1**

    (f) Calculate the mass of 0.5980 moles of sugar.

(1 mark) **KA1**

    (g) Calculate the number of moles of sugar molecules in a 4.2 g sample.

(1 mark) **KA1**

25. Magnesium has three naturally occurring stable isotopes: magnesium-24, magnesium-25, and magnesium-26.

Isotope	Atomic Mass	% Abundance on Earth
magnesium-24	23.985	79.0
magnesium-25	24.986	10.0
magnesium-26	25.983	11.0

Calculate the relative atomic mass of magnesium.

(3 marks) **KA1**

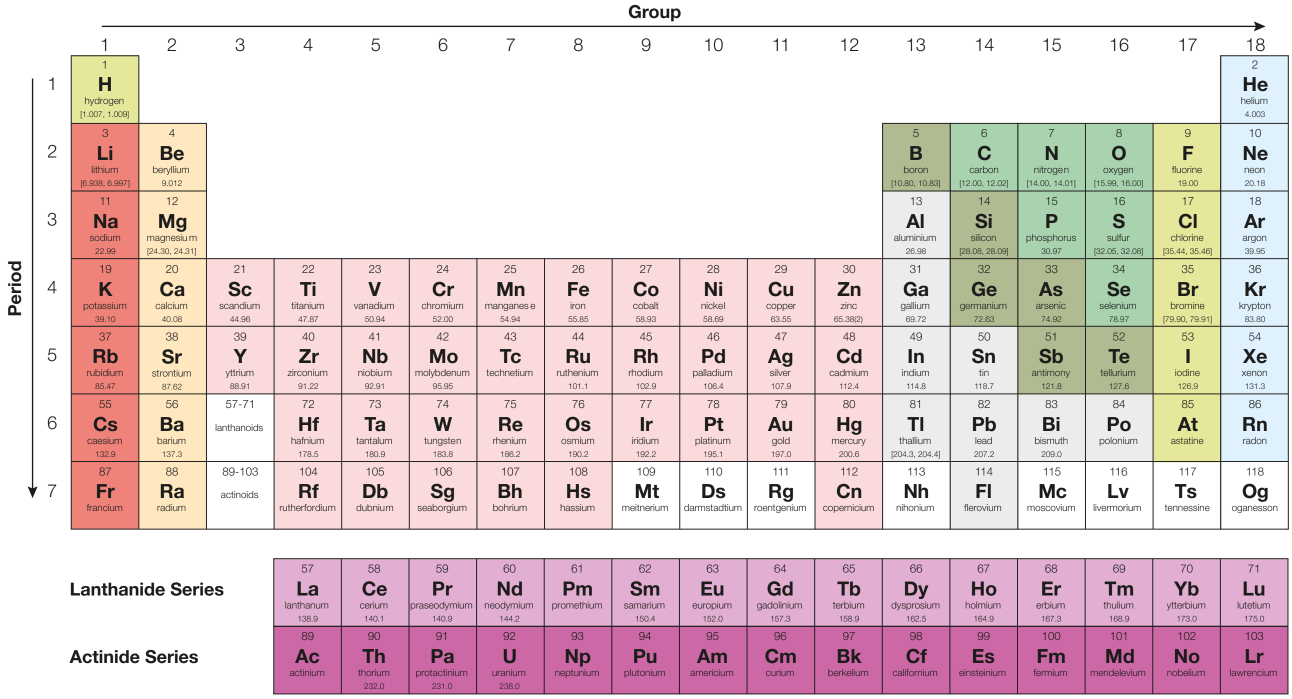
1.4 The periodic table

Science understanding

In the modern periodic table, elements are arranged in order of increasing atomic number, and display periodic trends in their properties.

- Identify the position of an atom in the periodic table given its electron configuration.
- Identify the *s*, *p*, *d*, and *f* blocks of the periodic table.

The **periodic table** displays the 118 known elements in order of increasing atomic number.



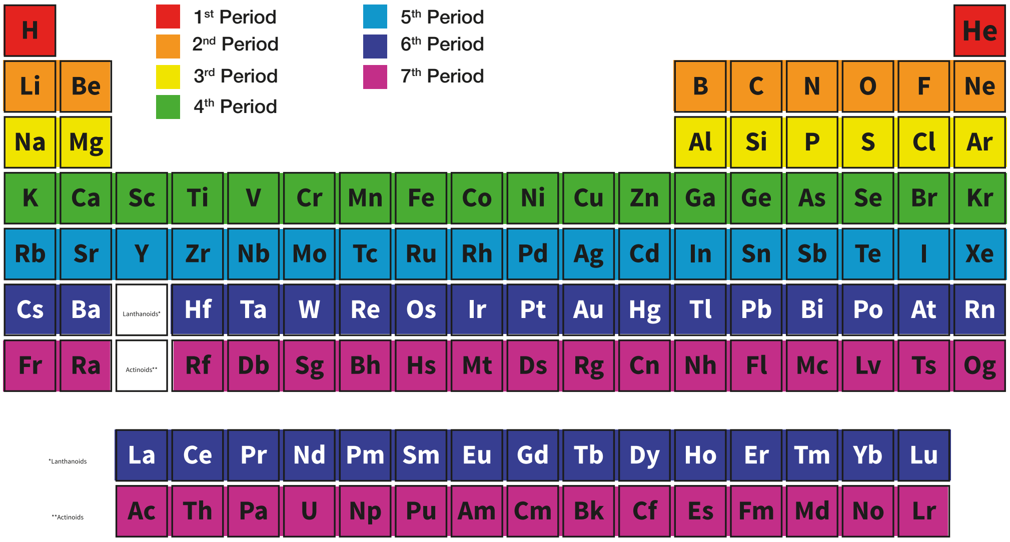
Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** hydrogen [1.007, 1.009]																	2 **He** helium 4.003
2	3 **Li** lithium [6.938, 6.997]	4 **Be** beryllium 9.012											5 **B** boron [10.80, 10.83]	6 **C** carbon [12.00, 12.02]	7 **N** nitrogen [14.00, 14.01]	8 **O** oxygen [15.99, 16.00]	9 **F** fluorine 19.00	10 **Ne** neon 20.18
3	11 **Na** sodium 22.99	12 **Mg** magnesium [24.30, 24.31]											13 **Al** aluminium 26.98	14 **Si** silicon [28.08, 28.09]	15 **P** phosphorus 30.97	16 **S** sulfur [32.05, 32.08]	17 **Cl** chlorine [35.44, 35.46]	18 **Ar** argon 39.95
4	19 **K** potassium 39.10	20 **Ca** calcium 40.08	21 **Sc** scandium 44.96	22 **Ti** titanium 47.87	23 **V** vanadium 50.94	24 **Cr** chromium 52.00	25 **Mn** manganese 54.94	26 **Fe** iron 55.85	27 **Co** cobalt 58.93	28 **Ni** nickel 58.69	29 **Cu** copper 63.55	30 **Zn** zinc 65.38(2)	31 **Ga** gallium 69.72	32 **Ge** germanium 72.63	33 **As** arsenic 74.92	34 **Se** selenium 78.97	35 **Br** bromine [79.90, 79.91]	36 **Kr** krypton 83.80
5	37 **Rb** rubidium 85.47	38 **Sr** strontium 87.62	39 **Y** yttrium 88.91	40 **Zr** zirconium 91.22	41 **Nb** niobium 92.91	42 **Mo** molybdenum 95.95	43 **Tc** technetium	44 **Ru** ruthenium 101.1	45 **Rh** rhodium 102.9	46 **Pd** palladium 106.4	47 **Ag** silver 107.9	48 **Cd** cadmium 112.4	49 **In** indium 114.8	50 **Sn** tin 118.7	51 **Sb** antimony 121.8	52 **Te** tellurium 127.6	53 **I** iodine 126.9	54 **Xe** xenon 131.3
6	55 **Cs** caesium 132.9	56 **Ba** barium 137.3	57-71 lanthanoids	72 **Hf** hafnium 178.5	73 **Ta** tantalum 180.9	74 **W** tungsten 183.8	75 **Re** rhenium 186.2	76 **Os** osmium 190.2	77 **Ir** iridium 192.2	78 **Pt** platinum 195.1	79 **Au** gold 197.0	80 **Hg** mercury 200.6	81 **Tl** thallium [204.3, 204.4]	82 **Pb** lead 207.2	83 **Bi** bismuth 209.0	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
7	87 **Fr** francium	88 **Ra** radium	89-103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **Nh** nihonium	114 **Fl** flerovium	115 **Mc** moscovium	116 **Lv** livermorium	117 **Ts** tennessine	118 **Og** oganesson

Lanthanide Series	57 **La** lanthanum 138.9	58 **Ce** cerium 140.1	59 **Pr** praseodymium 140.9	60 **Nd** neodymium 144.2	61 **Pm** promethium	62 **Sm** samarium 150.4	63 **Eu** europium 152.0	64 **Gd** gadolinium 157.3	65 **Tb** terbium 158.9	66 **Dy** dysprosium 162.5	67 **Ho** holmium 164.9	68 **Er** erbium 167.3	69 **Tm** thulium 168.9	70 **Yb** ytterbium 173.0	71 **Lu** lutetium 175.0
Actinide Series	89 **Ac** actinium	90 **Th** thorium 232.0	91 **Pa** protactinium 231.0	92 **U** uranium 238.0	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Figure 1.38: The modern periodic table

Periods

Each row is called a **period** and represents one main electron shell. A new row is started when the highest energy electron exists in a higher main shell.



🟥 1st Period
🟧 2nd Period
🟨 3rd Period
🟩 4th Period
🟦 5th Period
🟦 6th Period
🟪 7th Period

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	Lanthanoids*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Actinoids**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

*Lanthanoids	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
**Actinoids	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Figure 1.39: Periodic table showing the periods

The period in which an element appears can be determined from its main-shell electron configuration or from its subshell electron configuration.

Example

Element 15, phosphorus, has 15 electrons.

Using Main-Shell Electron Configuration

The main-shell electron configuration is 2,8,5.

There are *three* numbers in the main-shell electron configuration.

Phosphorus appears in *period 3*.

Using Subshell Electron Configuration

The electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$.

The highest occupied *s* subshell ($3s^2$) is in the *third* main shell.

Phosphorus appears in *period 3*.

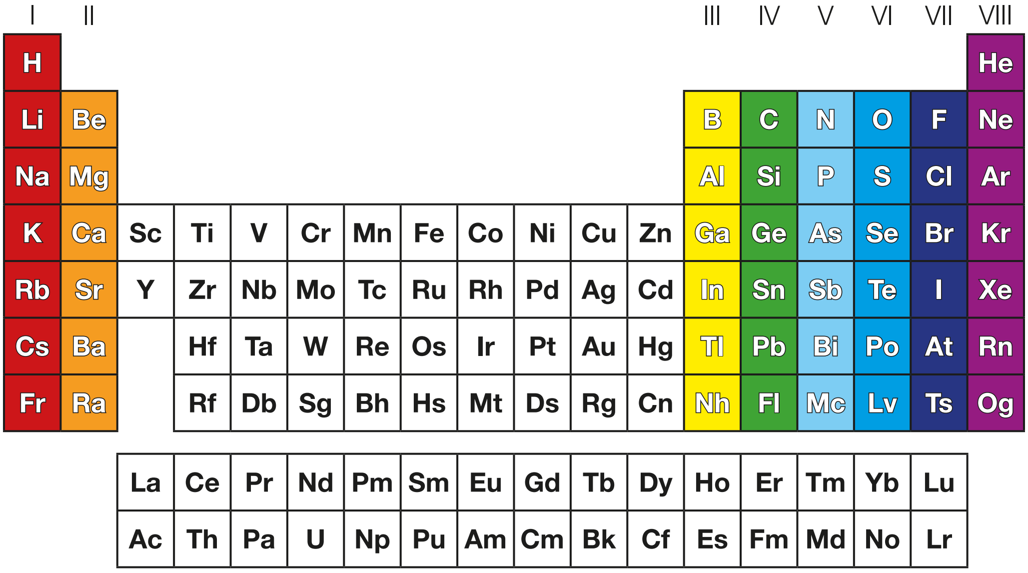
Groups

Each column on the periodic table is called a **group**. Groups are often labelled using Roman numerals.

Arabic Numerals	Roman Numerals
1	I
2	II
3	III
4	IV
5	V
6	VI
7	VII
8	VIII
9	IX

Arabic Numerals	Roman Numerals
10	X
11	XI
12	XII
13	XIII
14	XIV
15	XV
16	XVI
17	XVII
18	XVIII

The table is arranged such that the elements in each column have similar properties and electron configurations. There are 18 groups, and groups I–II and XIII–XVIII are called the **main groups**. The main groups are often numbered I–VIII. This means that, without context, a phrase like "group IV" is ambiguous, and could mean group IV (titanium, zirconium, hafnium, etc.) or *main* group IV (carbon, silicon, germanium, etc.)



I	II											III	IV	V	VI	VII	VIII
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Figure 1.40: Periodic table showing the main groups

The group in which an element appears can be determined from its main-shell electron configuration or from its subshell electron configuration.

Example

Element 15, phosphorus, has 15 electrons.

Using Main-Shell Electron Configuration

Its main-shell electron configuration is 2,8,5.

The final number in the main-shell electron configuration is *5*.

Phosphorus appears in *main group V*.

Using Subshell Electron Configuration

The electron configuration is $1s^22s^22p^63s^23p^3$.

The highest occupied subshell is $3p^3$.

Phosphorus appears in the *third* column in the *p-block*.

s-, p-, d-, and f-blocks

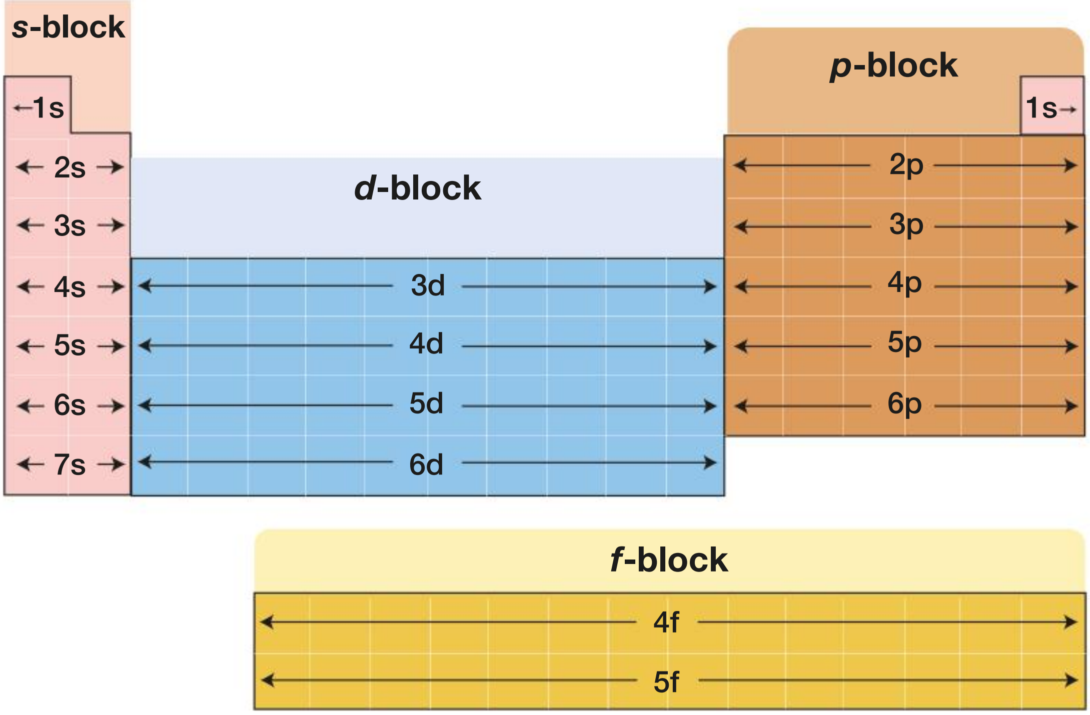


Figure 1.41: Periodic table showing the s-, p-, d-, and f-blocks

Groups I and II form the *s*-block. Elements in group I have electron configurations ending in s^1 and are called the **alkali metals**. Elements in group II have electron configurations ending in s^2 and are called the **alkaline earth metals**. Although helium has an electron configuration of $1s^2$, it is usually placed with group XVIII because it has very similar properties to the other elements in group XVIII.

Groups XIII-XVIII (main groups III-VIII) form the ***p*-block**. Elements in Groups XIII-XVI have electron configurations ending in p^1 to p^4. These elements have wildly varying properties but can be classified as the **post-transition metals**, the **metalloids**, and the **non-metals**. Elements in group XVII have electron configurations ending p^5 and are called the **halogens**. Elements in group XVIII have electron configurations ending in p^6 and are called the **noble gases**. The halogens and noble gases are also non-metals.

Groups III-XII form the ***d*-block**. These elements have electron configurations ending in d^1 to d^{10} and are known as the **transition metals**.

The elements with their highest energy electrons in *f* subshells (elements 57–71 and 89–103) have been placed in their own section at the bottom, the ***f*-block**, and the rest of the table has been concatenated. This is done to make the table easier to print and read, and because most of the *f*-block elements (the **lanthanides** and **actinides**) are extremely rare.

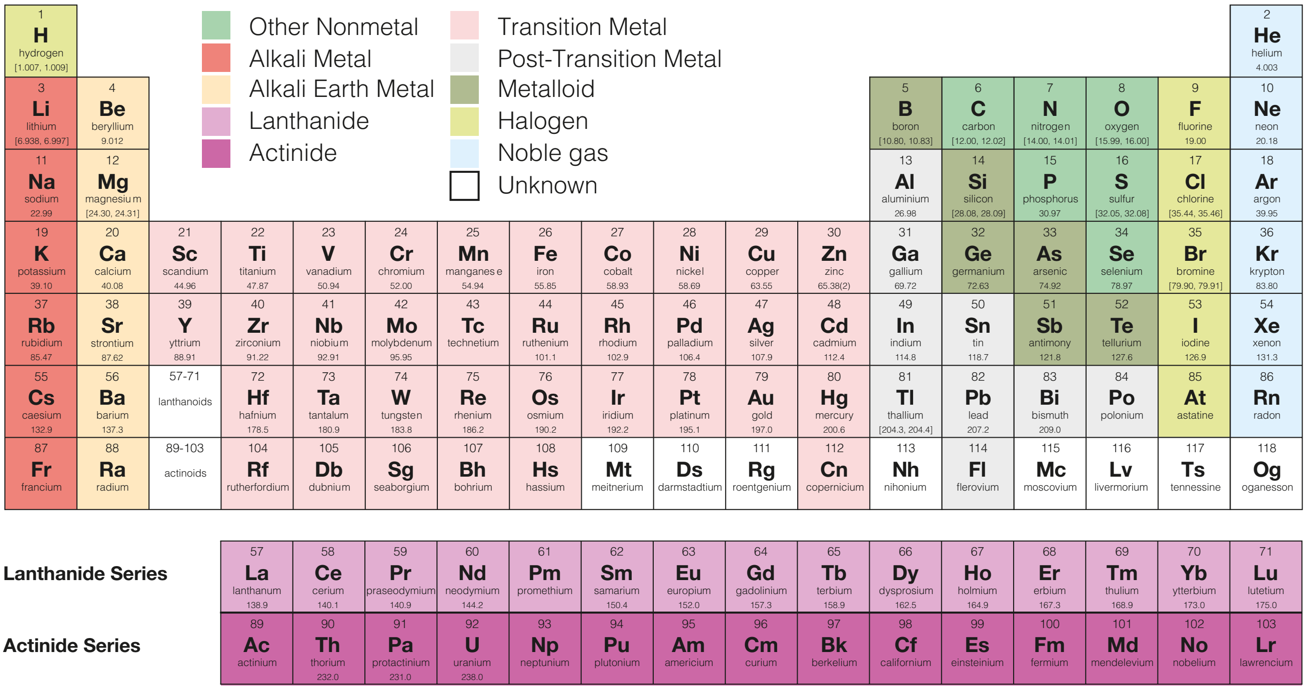


Figure 1.42: Periodic table showing the element classifications

Science understanding

The position of an element in the periodic table is related to its metallic or non-metallic character.

- Identify trends in atomic radii, valencies and electronegativities, across periods and down groups of the periodic table.

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Metallic and Non-Metallic Character

Most elements are classified as either **metals** or **non-metals** based on their properties, although some elements display both metallic and non-metallic properties and are known as **metalloids.** Metals are found on the left side of the table, and non-metals on the right, with the metalloids straddling the boundary. Hydrogen is an exception; it is the only non-metal found on the left side of the table.

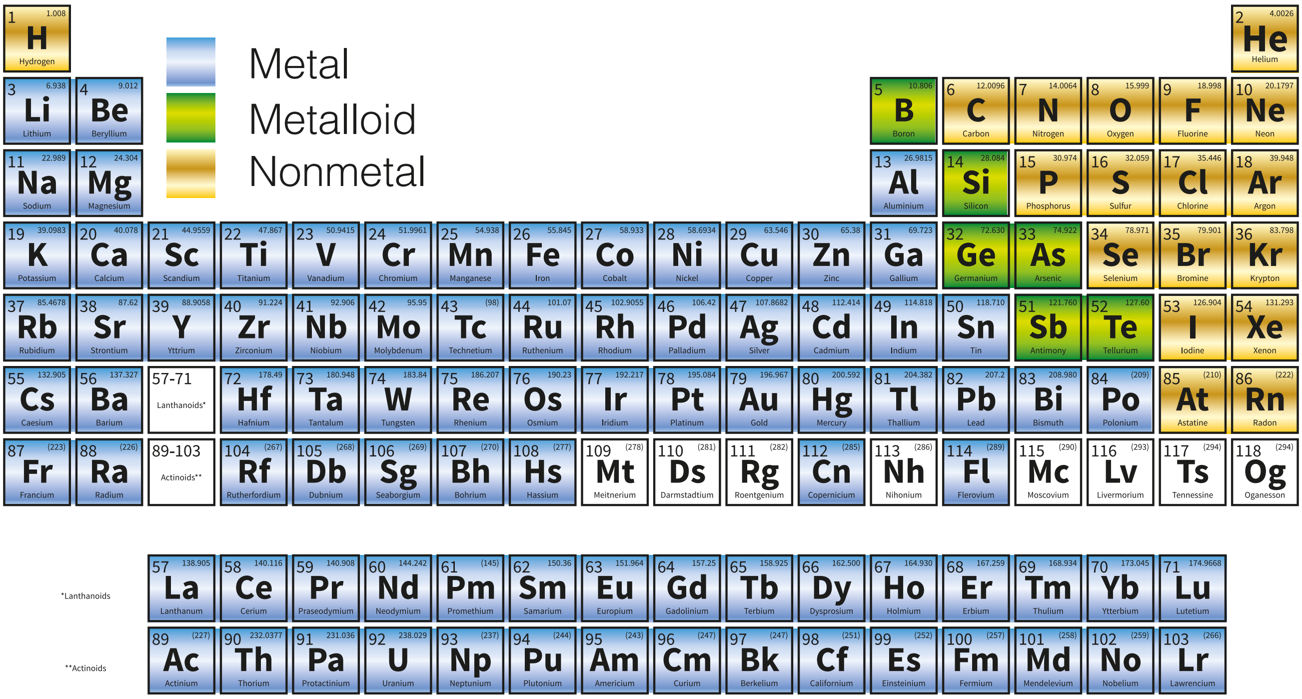


Figure 1.43: Periodic table showing classification of elements as metals, non-metals, or metalloids

Metallic Properties	Non-Metallic Properties
Strong/hard	Weak/soft
Lustrous (shiny)	Dull
Good conductors of electricity and heat	Poor conductors of electricity and heat
Ductile and malleable	Brittle (when solid)
High melting/boiling points	Low melting/boiling points
Low electronegativity	High electronegativity
Release electrons to form positive cations	Gain electrons to form negative anions

Valency

The **valency** of an atom is its capacity to form bonds with other atoms.

For main groups I–III and V–VII, the charge that an atom will take when forming an ionic bond can be predicted from the periodic table.

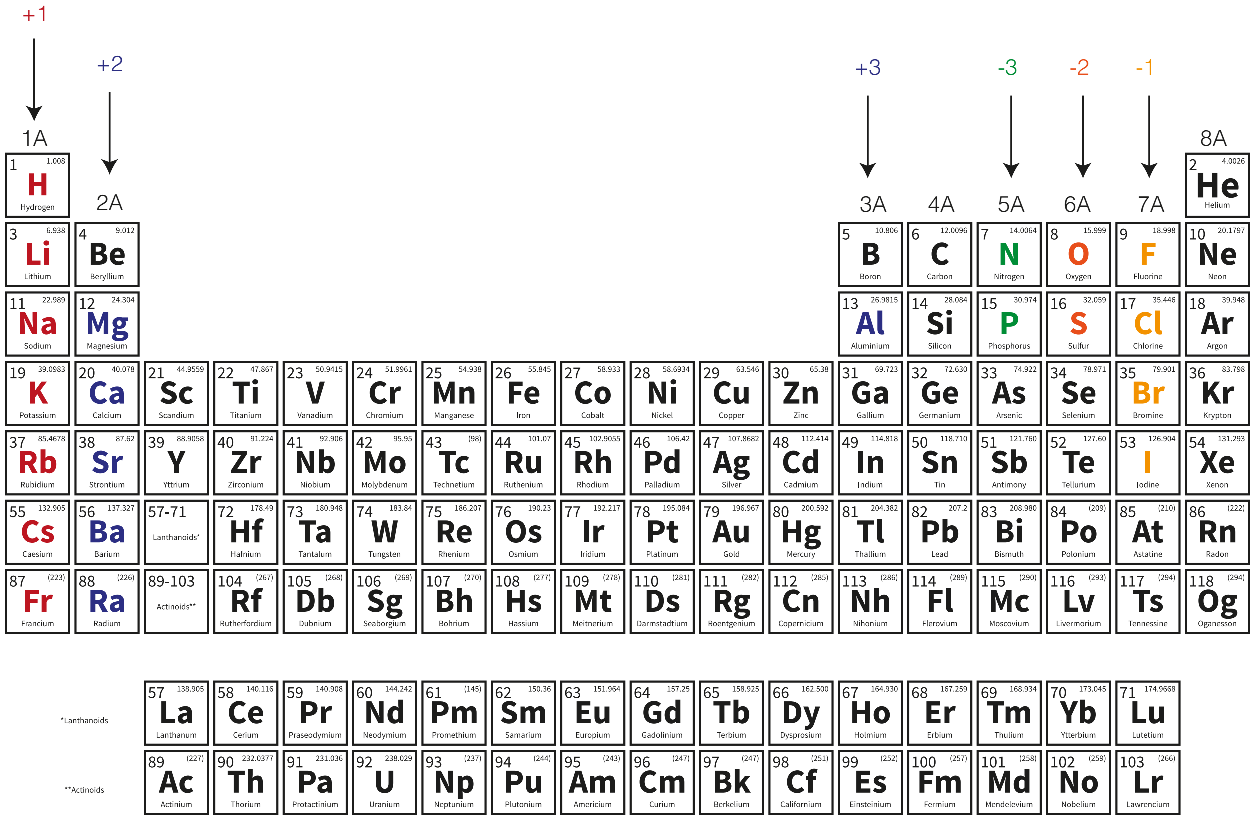


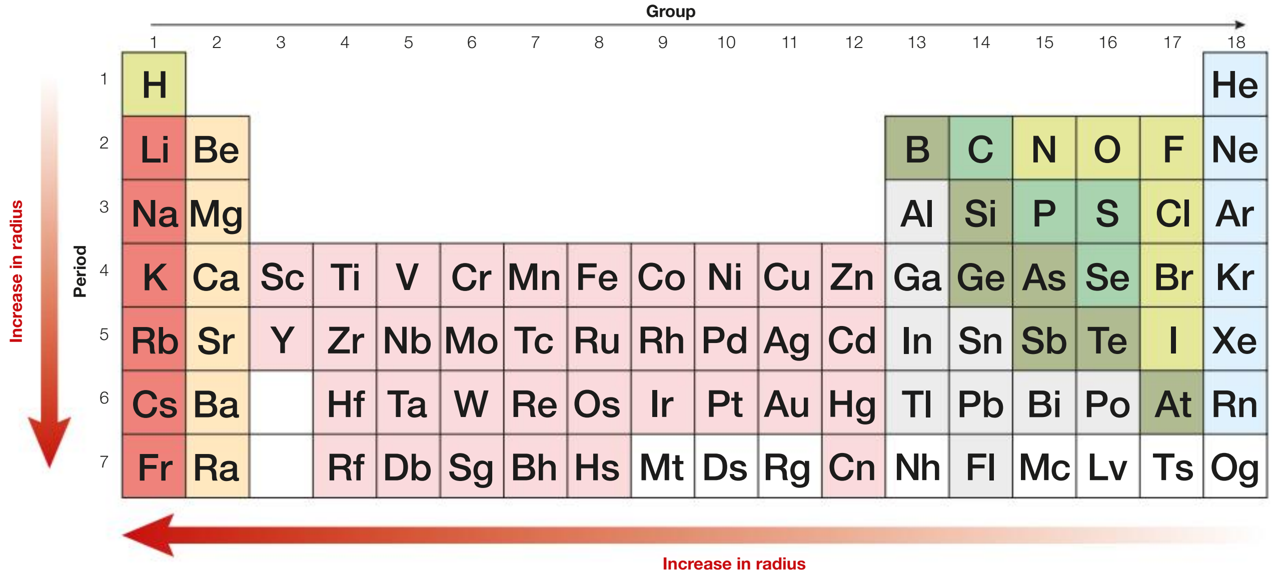
Figure 1.44: Periodic table showing predicted ionic charges

Atomic Radius

Atoms do not have sharply defined boundaries, but the relative sizes of atomic radii can be compared using the position of the outermost electrons.

Atomic radius increases as one travels down a group, as the electrons occupy higher-energy electron shells which are more distant from the nucleus.

Atomic radius decreases as one travels from left to right across a period. The additional electrons are added to the same main shell, and the additional protons present in the nucleus exert a greater attractive electrostatic force on the electrons which draws all the electrons closer to the nucleus.



Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

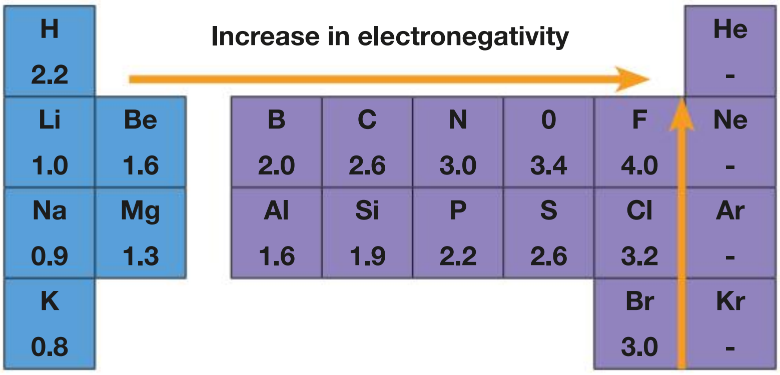
Figure 1.45: The trends in atomic radii

Electronegativity

The **electronegativity** of an element is a number representing the ability of an atom to attract an electron to itself. Electronegativity values are expressed on the Pauling scale using arbitrary units between 0 and 4. Caesium and francium have the lowest electronegativity (0.7), and fluorine has the highest (4.0).

Electronegativity decreases as one travels down a group for two reasons. Firstly, a larger atomic radius means the outermost shell is further away from the nucleus, and this weakens the attractive force. Secondly, there are more electrons between the nucleus and the outermost shell, and this creates a 'shielding' effect which also weakens the attractive force.

Electronegativity increases as one travels across a period because the reduced atomic radius means the outermost shell is closer to the nucleus, which strengthens the attractive force.



H **2.2**							**He** **-**
Li **1.0**	**Be** **1.6**	**B** **2.0**	**C** **2.6**	**N** **3.0**	**0** **3.4**	**F** **4.0**	**Ne** **-**
Na **0.9**	**Mg** **1.3**	**Al** **1.6**	**Si** **1.9**	**P** **2.2**	**S** **2.6**	**Cl** **3.2**	**Ar** **-**
K **0.8**						**Br** **3.0**	**Kr** **-**

Figure 1.46: Trends in electronegativity of the main group elements

The trends in atomic radius and electronegativity are inverse to one another. They are summarised in Figure 1.47.

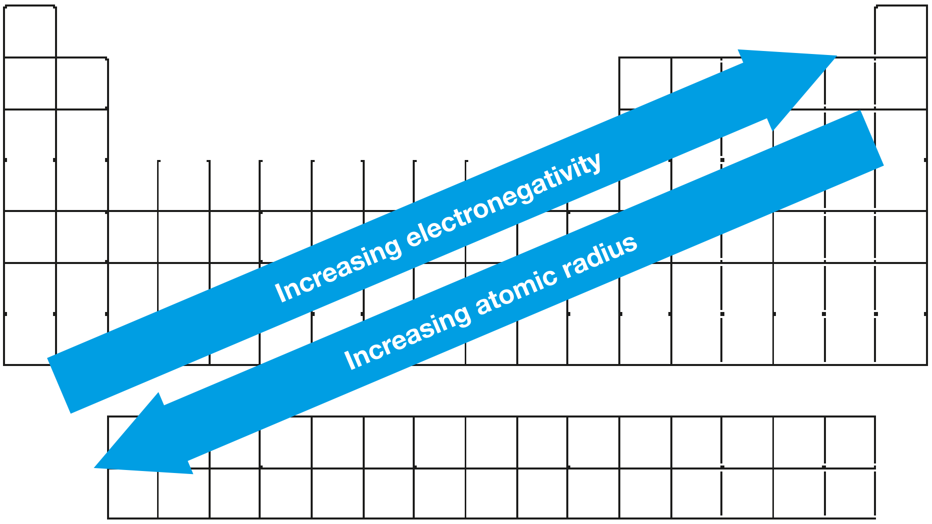


Figure 1.47: periodic table showing general trends for atomic radius and electronegativity

Questions

26. Complete the table below by classifying each element as a metal, non-metal, or metalloid.

Atomic Number	Symbol	Name	Classification
2	He	helium	non-metal
	Zn		
		silver	
	Ge		
		phosphorus	
5			
		arsenic	
19			
	Au		

(8 marks) **KA1**

27. Complete the table below by classifying each element as an alkali metal, alkaline earth metal, transition metal, post-transition metal, halogen, noble gas, lanthanide, or actinide. Additionally, state which block of the periodic table the element is found in.

Atomic Number	Symbol	Name	Classification	Block
54	Xe	xenon	noble gas	p-block
41				
12				
		lead		
	F			
		praseodymium		
	U			
37				
		krypton		

(8 marks) **KA1**

28. Use the electron configurations below to identify the period number, group number, and identity of each of the following neutral atoms below.

Electron Configuration	Period Number	Group Number	Element Name
$1s^22s^22p^2$			
$1s^22s^22p^63s^23p^64s^23d^6$			
$1s^22s^22p^63s^23p^6$			
$1s^22s^22p^63s^23p^64s^13d^5$			
$1s^22s^2$			
$1s^22s^22p^4$			
$1s^22s^22p^63s^1$			
$1s^22s^22p^63s^23p^5$			
$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$			

(8 marks) **KA1**

29. Fluorine and oxygen are both reactive non-metals.

    (a) "Fluorine is more electronegative than oxygen". Explain this statement.

....

....

.... *(1 mark)* **KA4**

    (b) State whether fluorine and oxygen have high or low electronegativity, compared with other elements.

.... *(1 mark)* **KA1**

    (c) State which element, fluorine or oxygen, has the higher electronegativity.

.... *(1 mark)* **KA1**

    (d) State which element, fluorine or oxygen, has the greater atomic radius.

.... *(1 mark)* **KA1**

    (e) Explain your answer to Question 28 (d).

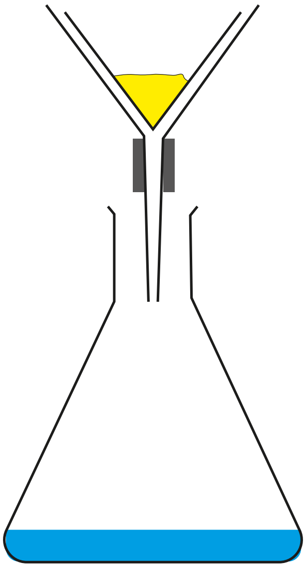
....

....

.... *(2 marks)* **KA1**

Summary Test 1: Materials and their Atoms

1. A mixture of sand and water can be separated using the apparatus shown below.



    (a) State the name of the separation technique shown.

.... *(1 mark)* **KA1**

    (b) Explain the separation technique shown.

....

....

.... *(2 marks)* **KA1**

Silicon dioxide is a major component of sand. It is a continuous covalent compound with a ratio of two oxygen atoms to each silicon atom (SiO_2).

(c) Draw and label a diagram of an atom of silicon-28.

(3 marks) **KA1**

(d) Classify silicon as a metal, non-metal, or metalloid.

.... *(1 mark)* **KA1**

(e) State the relative atomic mass of silicon.

.... *(1 mark)* **KA1**

(f) Calculate the relative formula mass of silicon dioxide (SiO_2).

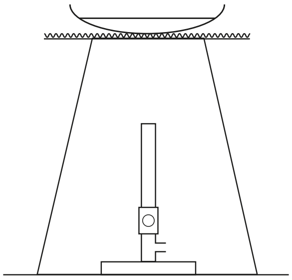
(1 mark) **KA1**

(g) State whether silicon atoms or oxygen atoms contribute more mass to the overall mass of silicon dioxide.

....

.... *(1 mark)* **KA1**

2. Solid sodium fluoride can be recovered from a solution of sodium fluoride in water using the apparatus shown below.



(a) State the name of the separation technique shown.

.... *(1 mark)* **KA1**

(b) Explain the separation technique shown.

....

....

.... *(2 marks)* **KA1**

(c) Explain why filtration cannot be used to separate the mixture.

....

....

.... *(2 marks)* **KA1**

(d) State the electron configuration of the sodium cation, Na^+.

.... *(1 mark)* **KA1**

(e) State the electron configuration of the fluoride anion, F^-.

.... *(1 mark)* **KA1**

3. Kettles intended for home use often use copper for their electrical wiring, nichrome for the heating element, and plastic for the handle.

(a) Identify the block on the periodic table that copper belongs to.

(1 mark) **KA1**

(b) State the property or properties of copper that make it suitable for use as wiring.

(1 mark) **KA1**

(c) Explain what would occur if the handle was also made from copper.

(2 marks) **KA1**

(d) State the property or properties of nichrome that make it suitable for use as a heating element.

(2 marks) **KA1**

(e) Explain what would occur if the electrical wiring in the cord was also made from nichrome.

(2 marks) **KA2**

4. Quantum dots (QDs) are particles of semiconducting material only a few nanometres in size. Similar to organic light-emitting diodes in OLED displays, they are used in QDLED displays to produce red, green, and blue light within each pixel.

Compared to OLED displays, QDLED displays can be brighter and show a wider variety of colours. QDLEDs can also produce thinner displays with increased physical flexibility.

(a) Explain why QDs are considered a nanomaterial.

(1 mark) **KA1**

(b) Describe the benefits to consumers of using QDLED displays, rather than older technologies.

(2 marks) **KA3**

Quantum dots can be made from a variety of compounds, including cadmium selenide. Selenium (atomic number 34) has five stable isotopes: ^{74}Se, ^{76}Se, ^{77}Se, ^{78}Se, and ^{80}Se.

(c) Using selenium as an example, define the term ‘isotope’.

(2 marks) **KA4**

(d) Selenium has a relative atomic mass of 78.97. Explain why the relative atomic mass is not a whole number.

(2 marks) **KA1**

5. In 1868, helium was discovered in the atmosphere of the Sun using absorption spectroscopy. A dark line was found in the yellow part of the visible spectrum, corresponding to a wavelength of 587.5 nanometres. Helium was not detected on Earth until 1881.

   (a) Draw and label a diagram of an atom of helium-2.

   (3 marks) **KA4**

   (b) Classify helium as a metal, non-metal, or metalloid.

   .... *(1 mark)* **KA1**

   (c) Describe the process of atomic absorption that produced the dark line.

   ....

   ....

   .... *(2 marks)* **KA1**

   (d) Explain how scientists in 1868 could be sure a new element had been detected.

   ....

   ....

   .... *(2 marks)* **KA2**

6. Neon lights operate by passing a high-voltage electrical current through a gas such as neon, exciting electrons in the neon atoms and causing atomic emission of light. Pure neon produces a reddish-orange colour; different colours of light are achieved by doping the neon with other elemental gases, and by using coloured glass in the construction of the tube.

   (a) State the main-shell electron configuration of neon.

   .... *(1 mark)* **KA1**

   (b) State the subshell electron configuration of neon.

   .... *(1 mark)* **KA1**

   (c) Explain why pure neon produces only one colour of light.

   ....

   ....

   .... *(2 marks)* **KA1**

(d) One potential electron configuration of excited neon is $1s^22s^22p^54p^1$. Explain how this electron configuration differs to the configuration of ground-state neon.

.. . .

.. . .

.. .. (2 marks) **KA1**

7. Depending on geographical location, each litre of seawater contains about 3.50 g of sodium chloride (NaCl).

(a) Calculate the relative formula mass of one formula unit of sodium chloride.

(1 mark) **KA1**

(b) Calculate the molar mass of sodium chloride.

(1 mark) **KA1**

(c) Calculate the number of moles of sodium chloride in 1 L of seawater.

(1 mark) **KA1**

(d) Calculate the number of sodium ions in 1 L of seawater.

(1 mark) **KA2**

Topic 2: Combining Atoms

2.1: Types of materials

Science understanding

Materials can be classified according to their structure and bonding into four types of substances.

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The formation of all solid materials results from **intramolecular** forces that form between atoms or ions on a microscopic level. The classes of intramolecular forces are **metallic**, **ionic**, **covalent (molecular or continuous)**. Intramolecular forces are often referred to as **primary** bonds.

Intramolecular Force	Element Type	Formed
Metallic	Metallic	Between positive metal cations and surrounded mobile sea of 'delocalised' electrons
Ionic	Metallic and Non-Metallic	Electrostatic attraction between a positive cation, and negative anion.
Covalent	Non-Metallic and Non-Metallic	Sharing of a pair of electrons between two atoms.

Table 2.01: Summary of intramolecular forces

Further detail about the formation of these bonds will be explored in Topic 2.2.

Solid materials are grouped into four types of substances: **metals**, **ionic compounds**, **covalent molecules** and **covalent networks.** Covalent molecules and covalent networks are both held together by covalent bonds, but are regarded as two distinct groups based on their unique and characteristic structures and properties.

Metallic solids such as magnesium, tin and copper are formed by the interaction of metal atoms. The regular **lattice** structure of metals is formed from the uniform distribution of atomic nuclei of metallic elements and a "**sea**" of **delocalised** valence electrons.

Figure 2.01: Metallic solids

Ionic solids such as sodium chloride and nickel oxide, are composed of interacting atoms of metallic and non-metallic elements. The geometrical structure of the ionic lattice results from the **electrostatic** attraction between positive **cations** and negative **anions,** which interact in a fixed ratio.

Figure 2.02: Ionic solid

Covalent molecular solids such as sucrose and iodine, are composed of discrete **molecules** which interact to form a solid material. Covalent bonding occurs when non-metallic atoms share valence electron pairs.

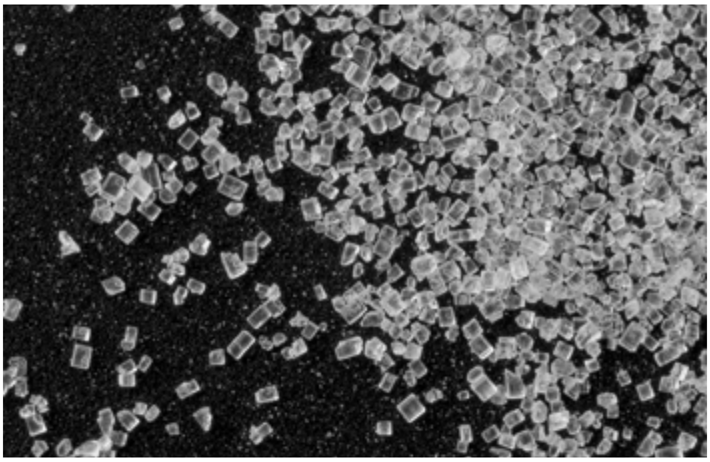

Figure 2.03: Covalent molecular solid: sucrose

Covalent network solids such as graphite and diamond, are composed of interacting atoms of non-metallic elements. The network lattice results from each atom forming covalent bonds with the nearest neighbouring atom at each valence electron site.

Figure 2.04: Covalent network solids: diamond, graphite and sand

Science understanding

Melting points can be used to classify materials into molecular and non-molecular lattices.

Melting Point

All materials have a characteristic **melting point** dependent on the strength of attraction between the particles present. The melting point is the temperature at which the solid material melts to become a liquid.

For solid materials, particles are tightly packed together in an orderly arrangement. The motion of individual atoms is restricted to vibrational motion around a fixed point. When heat energy is applied to a solid, the particles vibrate more rapidly. As this vibration increases, the organisation of the particles is disrupted, and the solid structure begins to break, resulting in the material starting to melt.

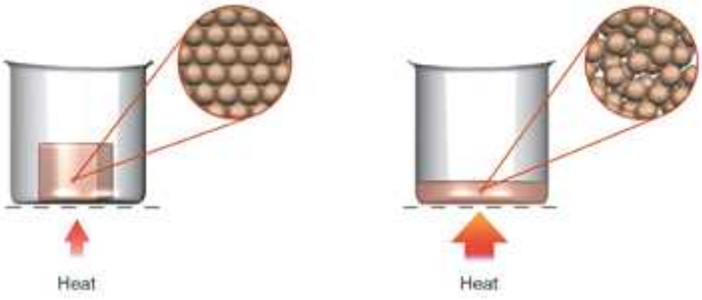

Figure 2.05: Transition from solid to liquid

For a solid to melt, the disruptive vibrations must exceed the degree of attractive intramolecular (between atoms) or **intermolecular** (between molecules) forces operating within the solid. Intramolecular bonding will be explored through this chapter, and *intermolecular bonding will be covered later in Topic 3*. Table 2.02 summarises the melting points of the different classes of solid materials.

Bonding Type	Particles	Bonding Strength	Example	Melting Point (°C)
Covalent network	Electronegative atoms	Very High	C (diamond) SiO_2	3500 1600
Metallic	Electropositive atoms	High	Ag Al	1064 660
Ionic	Ions	Moderate-High	NaCl KF	801 857
Covalent molecular	Molecules	Low	I_2 CCl_4	114 -23

Table 2.02: Melting Points of Solids

Science understanding

Electrical conductivity of non-molecular materials provides evidence for three types of primary bonding: metallic, ionic, and covalent.

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Electrical conductivity is a physical property that can be used to classify the intramolecular force present in a material. Electrical conductivity is the measure of the ease at which an electric **charge** can pass through a material. It involves the flow, or 'current', of free electrons or ions through a solid, molten or aqueous medium.

Delocalised electrons within a metallic lattice allow the flow of charge. In a similar way, ions in molten ionic liquids or aqueous ionic solutions can carry charge.

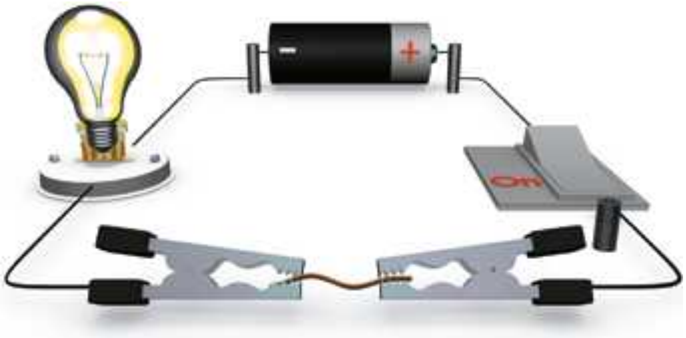

Figure 2.06: Electrical conductivity in metals

When ionic materials are in solid form, ions are locked into fixed positions. Without the presence of mobile ions, the material cannot carry charge. For this reason, covalent molecular solids and covalent network solids generally do not exhibit electrical conductivity, either in pure form or when dissolved in water.

Graphite, a covalent network lattice, is an exception; its layered structure includes some delocalised electrons between the layers. These electrons allow the movement of charge through the material, resulting in a material of excellent conductivity. Electrical conductivity based on a material's intramolecular bonding is summarised in Table 2.03.

Bonding	Conductivity		
	Solid	Molten	Aqueous Solution
Covalent network	No*	No	-
Metallic	Yes	Yes	-
Ionic	No	Yes	Yes
Covalent molecular	No	No	No

Table 2.03: Electrical Conductivity of Materials

*Yes, in rare substances with delocalised electrons (i.e., graphite)

This physical property is explained further in Subtopic 2.2.

2.2: Bonding between atoms

Science understanding

> The formation of bonds between atoms results in stable valence-shell configurations.
>
>

Octet Rule

Many atoms adhere to the **octet rule**, exhibiting a preference to attain eight **valence shell** electrons. This stable arrangement is either comprised of the atom’s own electrons, electrons that are donated by another atom, or electrons shared with another atom. An atom will usually continue to form bonds until an octet of valence electrons is achieved. Group VIII elements (noble gases) possess a complete octet in their atomic form and therefore are described as non-reactive, rarely forming compounds. Noble gases are the only group of elements to have atoms with a full octet and zero charge when unbonded.

Figure 2.07: Group VIII elements fulfil the octet rule

When valence electrons of atoms interact, the atom’s **electron configuration** changes. Metals release electrons when metallic bonds form, resulting in delocalised electrons and positively charged ions, known as cations. Figure 2.08 shows the formation of a lithium cation when it releases one electron to have a complete valence shell.

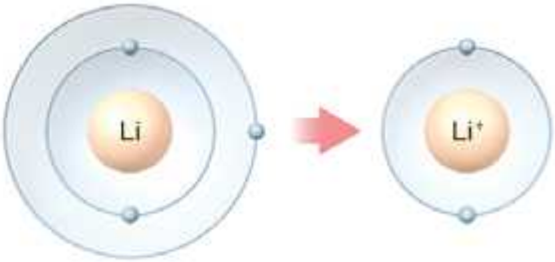

Figure 2.08: Formation of a lithium ion

Negatively charged ions, known as anions, form when non-metals gain electrons. Figure 2.09 depicts the formation of a fluoride anion when it gains one electron to complete its octet.

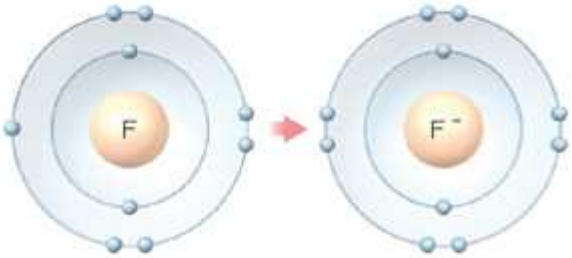

Figure 2.09: Formation of a fluoride ion

The group number of an element indicates the number of valence electrons that will participate in chemical reactions. From this, the number of bonds an atom will form can be determined; this is called the element’s **valency**. It also indicates the **magnitude** of the charge of ion that will form if the atom releases or gains electrons to complete the valence shell occupancy.

Group Number	Number of Valence Electrons	Charge of ion formed	Electrons
I	1	1+	Lost
II	2	2+	Lost
III	3	3+	Lost
IV	4	NA	NA
V	5	3–	Gained
VI	6	2–	Gained
VII	7	1–	Gained
VIII	8	NA	NA

Table 2.04: Valence electrons and charge of ions by group.

Questions

1. Group 1 metals such as potassium are highly reactive, requiring great care when being handled. They need to be stored under oil to avoid the metal coming into contact with moisture in the atmosphere.

   (a) State the valency of potassium.

   .... *(1 mark)* **KA2**

   (b) Describe the formation of the potassium ion.

   ....

   ....

   .... *(2 marks)* **KA1**

   (c) Write the formula of the potassium ion.

   .... *(1 mark)* **KA2**

2. Atom X has an electron configuration of $1s^2 2s^2 2p^4$.

   (a) State the number of electrons in Atom X's valence shell.

   .... *(1 mark)* **KA2**

   (b) State the name of this atom.

   .... *(1 mark)* **KA2**

   (c) State the group number of this atom.

   .... *(1 mark)* **KA2**

   (d) When Atom X becomes an ion, its formula is X^{2-}. Describe the formation of this ion.

   ....

   ....

   .... *(2 marks)* **KA1**

   (e) Write the electron configuration of X^{2-} ion.

   .... *(2 marks)* **KA1**

Science understanding

Energy is released when bonds are formed. Energy is required to break bonds.

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Bond Formation

A bond is formed when atoms combine to complete valence shells and stabilise their electron configurations. **Bond formation** involves the release of energy. This released energy usually takes the form of heat, resulting in a negative **enthalpy change**. The stored chemical energy, or heat content of the system, is referred to as enthalpy. *Enthalpy change is explained in further detail in Topic 4.4.*

Bond formation always releases energy. However, energy is required to rearrange bonds and can only occur when reactants possess energy equal to or greater than that of the **activation energy**. Activation energy is the minimum amount of energy that must be provided to compounds to result in a chemical reaction. The activation energy (E_a) of a reaction is measured in joules per mole (J/mol).

It is important to consider that when chemical reactions reaction take place, two separate steps are involved. The bonds of the reactants must first be broken, a process that requires energy. When new species are produced as a result of a reaction, the formation of these new bonds releases energy. The new arrangement of bonds does not have the same total energy as the bonds in the reactants.

Exothermic and Endothermic Reactions

In instances where the energy released when forming bonds of new products is less than the energy required to break bonds in reactant particles, the energy of the products will be higher than that of the reactants. The reaction is deemed **endothermic**, having absorbed energy from its surroundings. This relationship can be summarised using an **energy profile diagram**. An energy profile is a diagram that depicts the change in **chemical potential energy (enthalpy)**, referred to as the energy pathway as the reaction proceeds from reactants to product.

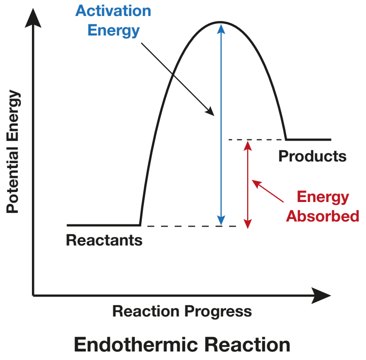

Figure 2.10: Energy profile diagram for an endothermic reaction.

When metals release electrons, the process is endothermic as energy is needed to remove the electron from the atom.

$$Na \rightarrow Na^+ + e^- \quad \Delta H = +496 \text{ kJ/mol}$$

For **exothermic** reactions, a greater quantity of energy is released when forming new bonds than what is absorbed to break the bonds of the reactants. Therefore, the enthalpy of the products is lower than that of the reactants. The energetic pathway for an exothermic reaction is shown in the energy profile diagram in Figure 2.11 below.

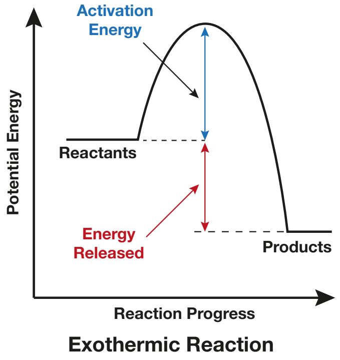

Figure 2.11: Energy profile diagram for an exothermic reaction.

Non-metals can gain electrons when combining with other atoms resulting in a release of energy as the atom completes its octet electron configuration. Gaining an electron is an exothermic process.

$$Cl + e^- \rightarrow Cl^- \quad \Delta H = -349 \text{ kJ/mol}$$

Questions

3. Fluorine is critical for the production of nuclear material for power plants and for the insulation of electric towers. As a non-metallic element, fluorine gains an electron to attain octet stability.

$$F + e^- \rightarrow F^- \quad \Delta H = -328 \text{ kJ/mol}$$

(a) The process of fluorine gaining an electron is an exothermic process. Explain this statement, referring to the equation above.

(2 marks) **KA1**

(b) Fluorine can react with hydrogen to produce hydrogen fluoride, or hydrofluoric acid via the following reaction:

$$F_2 + H_2 \rightarrow 2HF \quad \Delta H = -271 \text{ kJ/mol}$$

Depict the energy changes of the reactants to the products by drawing an energy profile diagram. State if the reaction is exothermic or endothermic and justify the response.

(6 marks) **KA4**

(c) Describe what is meant by the term activation energy. Use the reaction of fluorine and hydrogen to assist with this explanation.

(2 marks) **KA1**

Science understanding

Metallic bonding is the force of attraction between metal cations and their delocalised valence electrons.

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Metals

Metal materials are composed of bonded atoms from one or more metallic elements arranged in an orderly manner. Due to their useful set of properties, metals are valuable materials. Metals are **hard, malleable, ductile, lustrous**, have high melting and boiling points and are both electrical and thermal conductors. Non-metallic elements can be incorporated into metal materials in trace amounts, resulting in the formation of an **alloy**. Alloys have preferential properties over their base metal constituent.

Science as a human endeavour:

Advanced metal alloys

Advanced metal alloys have a vast array of applications in modern society, from construction materials to satellites. Creating new alloys with specific properties for intended purposes has been limited by researcher's understanding of what happens at the boundaries between the tiny crystalline grains within the metal's structure.

When metals are combined, the atoms of the secondary element may collect along these grain boundaries or spread out through the lattice. The alloy's properties are largely determined by the behaviour of these atoms.

Researchers at MIT have collaborated to develop a way to predict these properties using computer simulations and a machine-learning process. Such technology may facilitate the development of a wide variety of new applications for alloys in the future.

Metallic Bonding

Metals account for about three quarters of the 118 known elements on the periodic table, totalling 91 of the currently discovered elements. They are often valuable materials due to their properties and make up approximately 24% of the total mass of the planet.

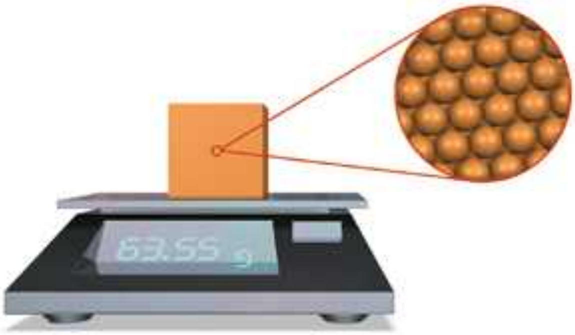

Figure 2.12: Structure of a metallic solid

Metallic bonds result in a formed material that is comprised of a tightly packed, dense and 3D regular lattice.

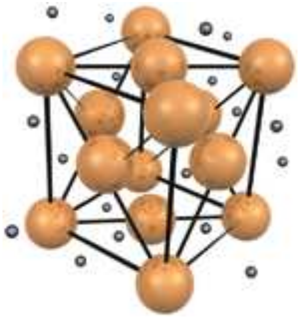

Figure 2.13: 3D metallic lattice

The **atomic radii** of elements increase down a period and from the right to left across a period. From exploration of this periodic trend, it can be concluded that metallic atoms have atomic radii greater than non-metallic elements of the same period.

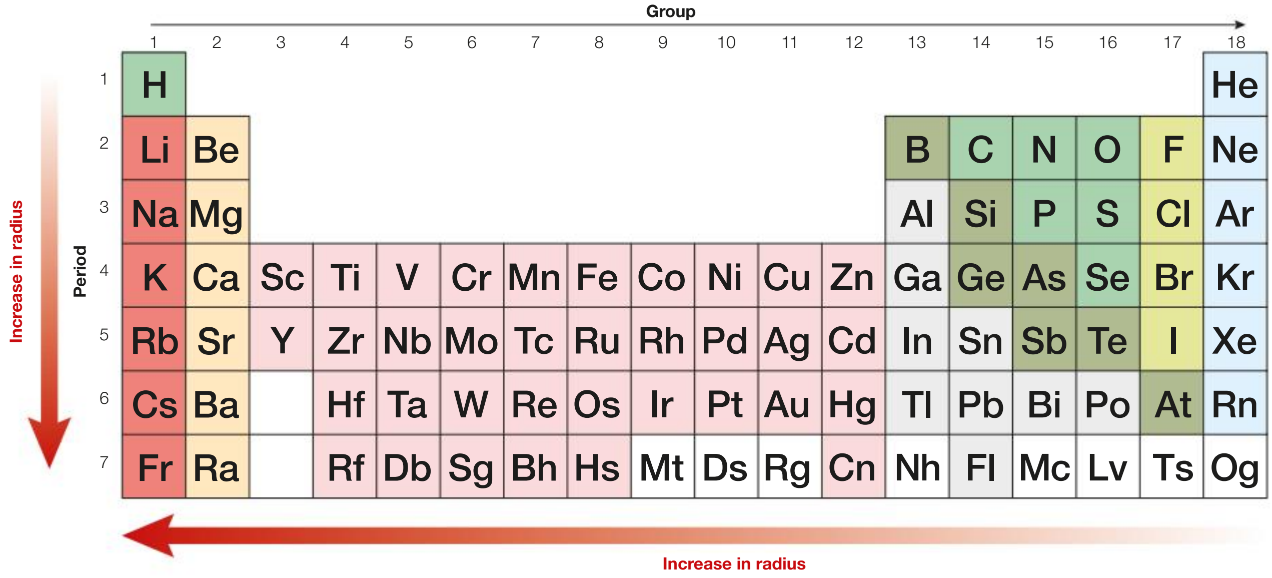

Figure 2.14: Trends in atomic radii

For atoms, this means that the negatively charged valence electrons are a further distance from the positively charged nuclei, resulting in a weaker attraction. These electrons are more easily lost in chemical reactions due to this weaker association.

The valence electrons of each metallic atom are shared by more than one neighbouring atom. When atoms come together, the electrons in the valence shell shares space with the corresponding electrons of the neighbouring atom. The valence shells of these adjacent atoms overlap, allowing electrons to move freely within these valence shells, detached from the electron's original atom. These electrons are referred to as delocalised electrons.

The metal atoms are held together by a "sea" of delocalised electrons. This results in a lattice of positive metal ions that the delocalised electrons move through. Although the positive nuclei repel each other, strong electrostatic attractions between the positive metal ions and delocalised electrons hold the metallic lattice together. The lattice is formed in layers and this model of metallic bonding explains the unique physical properties of metals.

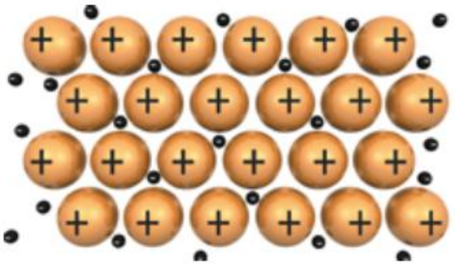

Figure 2.15: Positive nuclei and the "sea" of delocalised electrons

Science as a human endeavour:

Metals in medicine

Metals have had therapeutic applications throughout history. Written accounts date back to 1500BC detailing the use of copper to assist with inflammation and iron in the treatment of anaemia.

Modern developments have seen the application of metal complexes to treat cancer by direct interaction with the cancerous cell's DNA. The positive charge on metal cations can interact with the negative charge of the phosphate backbone of DNA. This interaction can result in oxidative damage to the DNA preventing further replication. Such agents encompass metallodrugs in which a metal complex is itself the active agent.

The relationship between metals and biological molecules provides infinite opportunities that are the subject of much research and development.

Questions

4. Titanium, a transition metal has many uses including the construction of replacement hip sockets.

   (a) Describe how titanium atoms interact to form a metallic solid.

   (2 marks) **KA1**

   (b) Explain what is meant by the term ‘delocalised’ electron.

   (2 marks) **KA1**

   (c) In reference to atomic radius, predict if the metallic bonding present in titanium would be stronger than the bonding present in zinc.

   (2 marks) **KA2**

Science understanding

The physical properties of metallic elements can be explained using the model for metallic bonding.

Explain the melting and boiling points, and electrical conductivities of metallic elements.

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Physical properties are properties of a material that can be observed or measured without changing the composition of the material.

Hardness

Hardness is a material’s ability to resist **deformation** or **compression**. It is commonly determined by tests which measures a material’s **resistance** to indentation.

Figure 2.16: Hardness of a metallic solid

Metals are hard materials, as the strong electrostatic forces of attraction between the positively charged nuclei and the “sea” of delocalised electrons hold the metallic lattice together. The atoms within a metallic solid are arranged in layers. The delocalised electrons within the structure create significant electron pressure. To compress them, pressures greater than this internal electron pressure are required. When a force is applied to compress the material, these layers may slide over one another. The harder the metal, the greater the force required to alter this structure.

Example

Alloy strength

In an alloy the atoms in the lattice structure are of different sizes. The difference in atomic size distorts the regular lattice structure. This results in alloys requiring a greater quantity of force to allow the layers to slide over one another, therefore alloys are often harder than the pure metals they are comprised of.

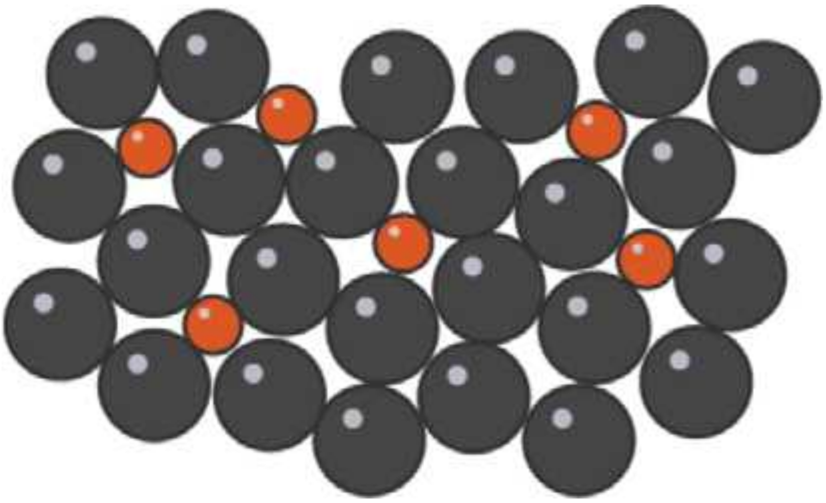


Figure 2.17: Lattice structure of a metal alloy

Malleability

Closely connected to the property of hardness is **malleability**. Malleability is a measure of a material's ability to withstand compressive stress, taking on a new shape without breakage. When a force is applied the positively charged nuclei are forced to slide over one another, however are held together by the electrostatic attraction between them and the delocalised electrons.

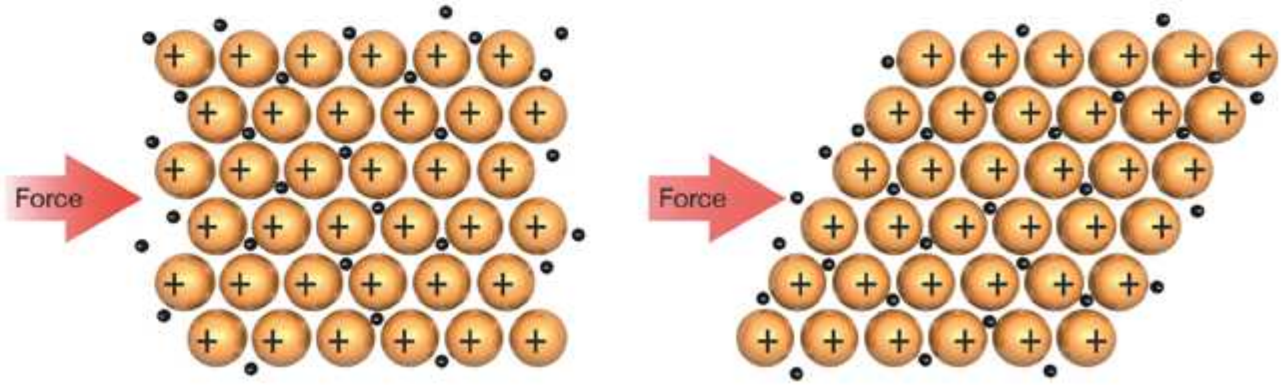


Figure 2.18: Malleability of a metal

The degree of deformation is dependent on the malleability of the material and the intensity of the force applied. The deformity may be temporary when a relatively small stress is applied to the material, or permanent where the atoms roll over each other into new positions. The strong intramolecular forces that hold them together remain unbroken. The malleability of a metallic lattice can be enhanced by heating the material.

Ductility

Ductility is a material's ability to resist **tensile** stress, a force which pulls a material in two opposing directions. Metals are ductile, which means they can withstand high tensile stress to be drawn into wires. When such stress is applied the attraction between positively charged nuclei and delocalised electrons are stretched but not broken due to their strong association.

Electrical Conductivity

Electrical conductivity is the measure of the ease at which an electric charge can pass through a material. Metals are electrical conductors as they have charged particles that are free to move. Delocalised electrons will move towards a positive electrode and away from a negative electrode in an electric circuit.

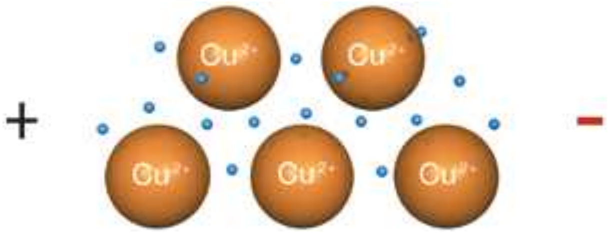


Figure 2.19: Electrical conductivity in metals (atomic level)

Example

Silver

Silver has the greatest electrical conductivity, with all other metals having their conductivity measured relative to that of silver. Due to this property, silver is often utilised as a component in electrical circuits and contacts.

Thermal Conductivity

Thermal conductivity refers to the transfer of heat through a material. Metals have high thermal conductivity. As heat energy is applied to a material the free delocalised electrons vibrate more rapidly, causing a greater number of spontaneous **collisions**. This transmits energy throughout the material distributing the heat rapidly.

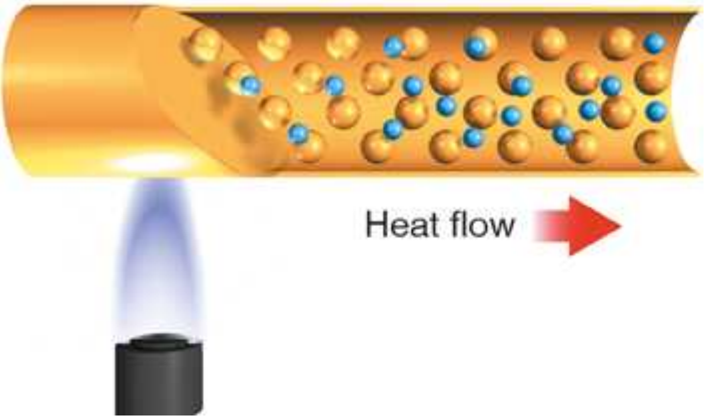


Figure 2.20: Thermal conductivity in metals

Density

Density describes the quantity of mass in a given volume. Metallic lattices result from the electrostatic attraction between delocalised electrons and positively charged nuclei, which are usually closely packed together. This results in different metals having a high, but varied density. The density of a metal depends on the mass of the metal ions, their radius and the way they are arranged in the lattice.

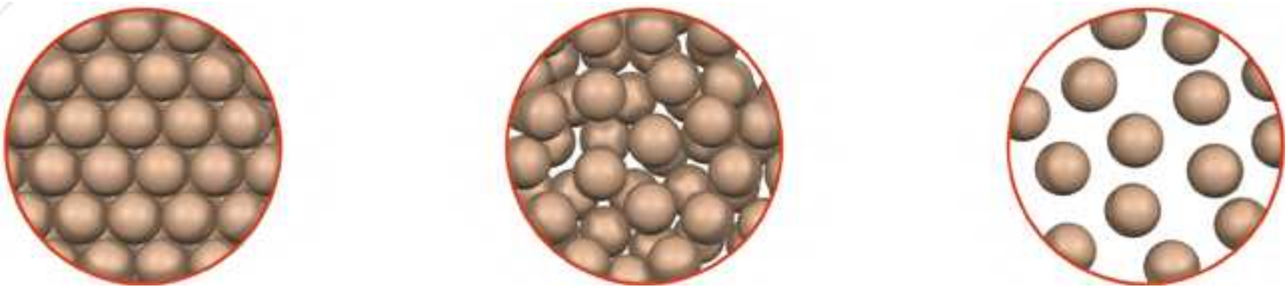


Figure 2.21: Cation "packing" in a metal lattice

Lustre

Lustre describes the reflective appearance of a material. Metals are lustrous materials. When light falls on the metal surface, photons of light are absorbed by the delocalised electrons. These electrons move from one energy level to a higher energy level. As the electron again drops back to a lower energy level, it emits energy in the form of photons. This movement of electrons between energy levels, causes electrons to vibrate at the atomic level. The wavelength of energy emitted, or reflected by the electrons results in a lustrous appearance. Different metals emit different wavelengths of energy.

Melting point

Melting point is the temperature at which the solid melts to become a liquid. The melting point of a solid is dependent on the strength of the bonding present. For metallic solids, strong electrostatic forces of attraction between positively charged nuclei and the "sea" of delocalised electrons hold the metallic lattice together. This explains the high melting points exhibited by metallic materials. When metallic bonds are provided with sufficient energy the cations will separate resulting in a molten compound.

Like density, the melting points of metals are high but varied. The strength of intramolecular forces is dependent on both the magnitude of charge and atomic radius of the metal cation involved. Metals with a greater **charge density** will have a higher melting point due to the greater electrostatic attraction between each positively charged nuclei and delocalised electrons. Metals of smaller atomic radius will have greater melting points as the positive nuclei are in closer proximity to the delocalised electrons, increasing the strength of electrostatic attraction.

Figure 2.22: Melting point of metals

Example

Magnesium has a melting point of 650°C, Aluminium has a slightly higher melting point of 660°C. Use the metallic bonding model to explain this slight difference.

Aluminium is a group III metal, possessing a 3+ cation when bonding. This is a greater charge magnitude than that of magnesium, a group II metal which produces a 2+ cation. The aluminium cation has greater electrostatic attraction for delocalised electrons increasing the amount of energy required to overcome the attraction to melt it.

Atomic radii increases moving from right to left across a period, therefore magnesium has a greater atomic radius than aluminium. This increases the distance of the positively charged nuclei from the delocalised electrons. This decreases the strength of the electrostatic attraction formed, decreasing the melting point for magnesium in comparison to aluminium.

Boiling point

The **boiling point** of a material is the temperature and pressure at which a liquid changes phase to a gas. As with melting points, the high boiling point of metallic materials is attributed to the strength of the electrostatic attraction between the positively charged nuclei and the "sea" of delocalised electrons.

Questions

5. Barium is a metallic element in group 2 of the periodic table. It has a melting point of 850°C and conducts electricity in solid state.

   (a) Describe the forces that hold the barium metallic lattice together.

....

....

....

.... (2 marks) **KA1**

(b) Describe how barium's high melting point can be explained in reference to its structure.

....

....

....

.... *(2 marks)* **KA1**

(c) Barium can also be flattened under compressive force. State the property that describes its ability to be deformed under pressure.

.... *(1 mark)* **KA2**

(d) Electrical conductivity is the measure of the ease with which an electric charge can pass through a material. In reference to barium's structure explain its conductivity.

....

....

....

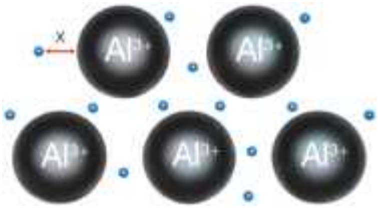
.... *(2 marks)* **KA1**

(e) Barium is often used in spark plug electrodes. State two physical properties that make it suitable for this use.

....

.... *(2 marks)* **KA2**

(f) Aluminium is a group III metal. The bonding present in the aluminium lattice is represented below:



Identify the attractive force labelled in the diagram above.

.... *(1 mark)* **KA2**

(g) Compare the atomic structure of the barium and aluminium to predict which metal has the higher melting point.

....

....

....

.... *(3 marks)* **KA1**

6. Graphite is a non-metallic substance that can be lustrous and conducts electricity and heat. It is not malleable, breaking when a force is applied.

(a) Identify the properties graphite shares with metals.

....

....

....

.... *(2 marks)* **KA2**

(b) Describe the inferences that can be made about the structure of graphite based on the properties it has in common with metals.

....

....

....

.... *(3 marks)* **KA1**

Science understanding

Valence electrons are transferred from a metallic atom to a non-metallic atom to form ions.

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Ions

Unbonded atoms are electrically neutral, containing an equal number of positively charged protons, and negatively charged electrons. Though neutral, atoms are unstable (with the exception of the Group VIII noble gases), possessing incomplete valence shells. To obtain their octet electron configuration and complete valence shell occupancy, atoms release or gain electrons. **Ions** are the charged particles that form through the gain or release of electrons.

Cations

Cations are positively charged ions that are formed when metallic atoms release valence electrons to non-metallic atoms. Energy is required to release an electron due to the attraction between the electron and positive nucleus. The minimum amount of energy required for this process is termed the **ionisation energy**. Ionisation energy varies for each element and demonstrates **periodicity**, increasing across a period and decreasing down a group. Metal atoms typically have relatively low ionisation energies.

When energy equal to the ionisation energy is available, electrons are lost. Because electrons are negatively charged, the metal atom gains an overall positive charge equal in magnitude to the number of electrons that are lost. This results in a positively charged particle.

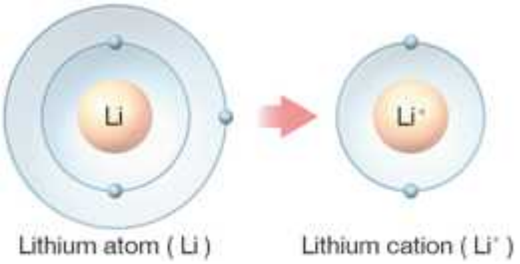


Figure 2.23: Removal of an electron from a lithium atom

This cation formation can also be expressed via an equation:

$$Li \rightarrow Li^+ + e^-$$

The magnitude of ionisation energy required increases for each subsequent electron lost.

The number of electrons lost is determined by the atom's group number. Atoms in group IIII have high ionisation energies and thus are unlikely to release or gain electrons to form ions. Recall that an atom's group number indicates the number of valence electrons.

Group Number	Number of Valence Electrons	Charge of cation formed
I	1	1+
II	2	2+
III	3	3+
IIII	4	Share electrons

Table 2.06: Formation of cations

To write the symbol of a cation, the element symbol is first recorded, followed by the charge in superscript. Cations are named according to their parental atom.

Parental Atom	Cation Symbol	Cation name
Sodium	Na^+	sodium cation
Magnesium	Mg^{2+}	magnesium cation
Aluminium	Al^{3+}	aluminium cation

Table 2.07: Cation symbol and naming conventions

Anions

Anions are negatively charged ions that are formed when non-metallic atoms gain electrons. Non-metallic atoms gain electrons as they typically have high ionisation energies. This means it is difficult to release their valence electrons. Noble gases have the highest ionisation energies, thus under normal conditions these atoms do not form ions. Non-metallic atoms have high electron affinity, therefore preference gaining electrons to complete their octet. When electrons are gained, negative charge equal in magnitude to the number of electrons is gained. This results in a negatively charged particle.

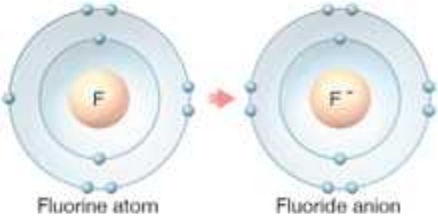


Figure 2.24: Fluoride anion formation

This anion formation can also be expressed via an equation:

$$F + e^- \rightarrow F^-$$

Anion formation also requires energy as the incoming electron must overcome the **repulsion** encountered between it and the existing valence electrons. Repulsion is experienced as particles with the same charge repel. If sufficient energy for the incoming electron is present, the electron successfully enters the valence shell, and a negative charge is acquired by the atom.

The number of electrons a non-metal will gain is predicted by the atom's group number. The difference between the octet and the group number indicates the number of electrons that will be gained.

Group Number	Number of Valence Electrons	Number of Electrons Gained	Charge of cation formed
V	5	3	3-
VI	6	2	2-
VII	7	1	1-
VIII	8	-	Do not form ions

Table 2.08: Formation of anions

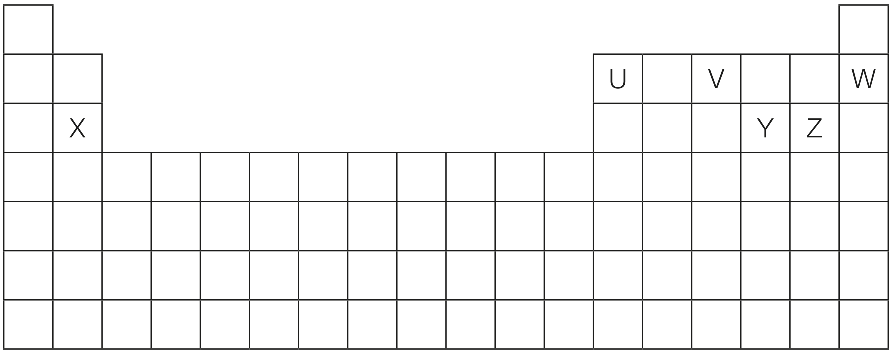
To write the symbol of an anion, the element symbol is first recorded, followed by the charge in superscript. Anions are named according to their parental atom, however the elemental suffix is removed and replaced with "*ide*".

Parental Atom	Cation Symbol	Cation name
Fluorine	F^-	fluoride anion
Oxygen	O^{2-}	oxide anion
Nitrogen	N^{3-}	nitride anion

Table 2.09: Anion symbol and naming conventions

Questions

7. The ion that an atom will form can be determined from the periodic table. Use the periodic table below to answer questions about elements U-Z.



(a) Write the formula of the ions formed by:

Atom	Formula and charge
V	
X	
Z	

(3 marks) **KA4**

(b) Classify the nature of the ion formed by the following atoms:

Atom	Anion / Cation
U	
X	
Y	

(3 marks) **KA4**

(c) Write a “T” in the periodic table to represent an atom that will form a T^- ion (an anion with an overall charge of 1–).

(1 marks) **KA2**

(d) Explain how cations acquire an overall positive charge.

(2 marks) **KA1**

(e) Atom “W” has electron configuration $1s^22s^22p^6$

(i) Complete the electron configuration for the following atoms:

Atom	Electron Configuration
V	
X	

(4 marks) **KA4**

(ii) Explain the similarity in electron configuration between “W” and the electron configurations formed when V and X become *ions*.

(3 marks) **KA1**

Science understanding

Ionic bonding is the force of attraction between the oppositely charged ions.

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Ionic Bonding

Metallic atoms tend to release electrons to non-metallic elements to form positively charged cations. Non-metallic atoms gain electrons from metallic atoms to form negatively charged anions. This transfer results in ions of opposing attractive electrical charges. Electrostatic attraction occurs between the two ions, releasing energy in the process, forming an ionic bond. This is a two-step process: ions are formed through the exchange of valence electrons, and then ionic bonds are the result of the electrostatic attraction. Ionic bonds facilitate the formation of **ionic compounds**.

Electrostatic attraction causes the cations and anions produced to form a lattice in which many ions are closely packed, arranged so that the cation-anion attraction is maximised, and the anion-anion or cation-cation repulsion is minimised. The strength of ionic attraction increases based on the degree of difference in the atoms' **electronegativity,** and the magnitude of charge of each ion. Ionic compounds are commonly referred to as **salts**.

Example

Sodium chloride is an ionic compound, specifically an ionic lattice. It is a solid which is formed from the interaction of a cation and an anion. With sufficient energy (equal or greater to the ionisation energy of sodium), an atom of sodium can release its valence electron resulting in the formation of the sodium cation (Na^+). This electron is accepted by chlorine, overcoming the repulsion resistance of its valence electrons to form the chloride anion (Cl^-). This can be depicted in an electron transfer diagram as shown below in Figure 2.25.

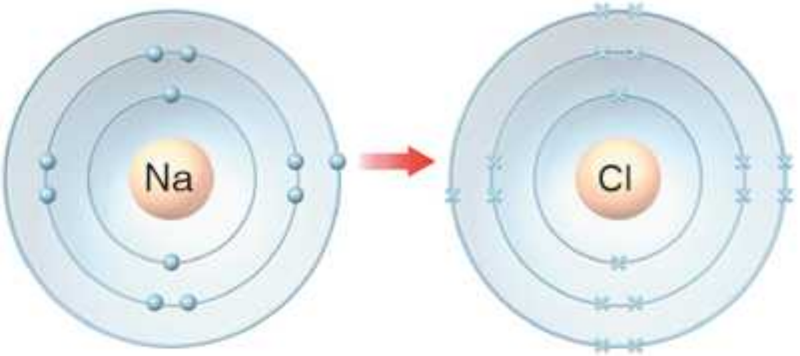

Figure 2.25: Transfer of electron from sodium to chlorine

An ionic bond forms between the sodium cation and chloride anion due to the electrostatic attraction between the oppositely charged ions. The ionic compound sodium chloride is formed.

$$Na^+ + Cl^- \rightarrow NaCl$$

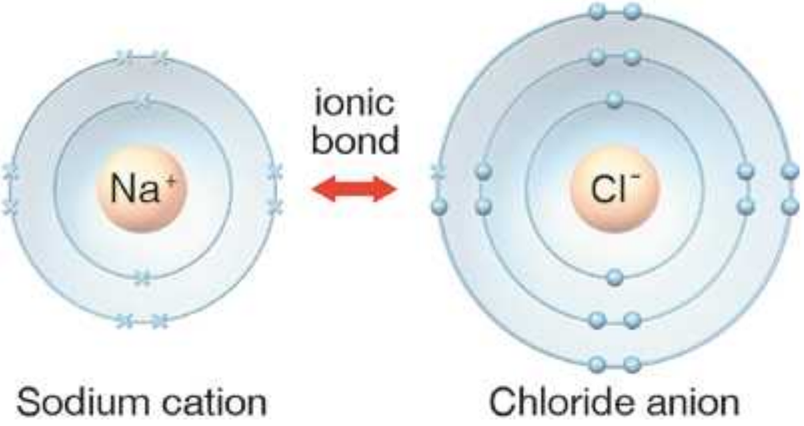

Figure 2.26: Ionic bond between sodium and chloride ions

Example

Magnesium chloride is an ionic compound formed by the interaction of magnesium cations (Mg^{2+}) and chloride anions (Cl^-). In this example there is a difference in charge magnitude of the cation and anion. Each chlorine atom can only accept one electron from the magnesium atom. As magnesium is a group II metal, it has two valence electrons to release. In these instances, one electron is transferred to one of two chlorine atoms.

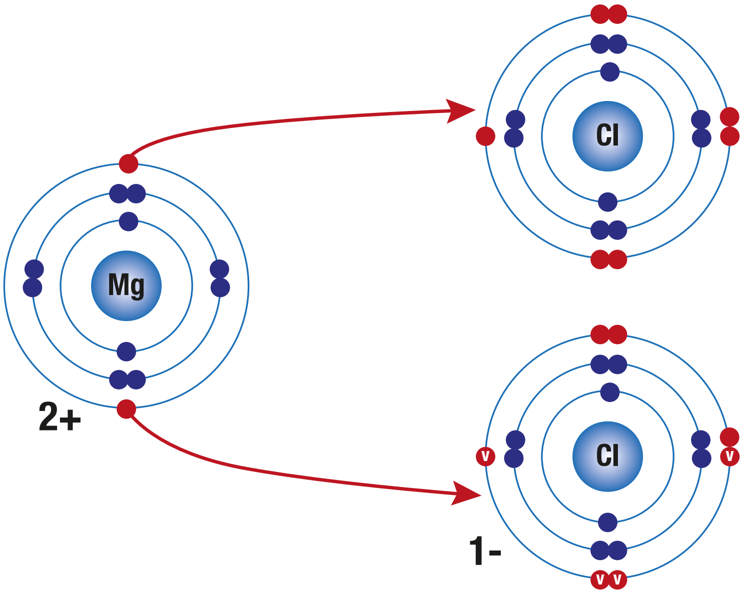

Figure 2.27: Formation of a magnesium ion and chloride ions

For each magnesium cation, two chloride anions are required to form magnesium chloride.

$$Mg^{2+} + Cl^- \rightarrow MgCl_2$$

Naming Ionic Compounds

Ionic compounds are named systematically.

1. The cation is first identified and named.
2. The anion is then identified and named, ensuring the suffix is replaced with "*ide*".

This process was outlined in the examples above for both sodium chloride and magnesium chloride. Further examples are outlined in the Table 2.10.

Cation	Anion	Ionic Compound
Lithium	Oxygen	Lithium oxide
Potassium	Bromine	Potassium bromide
Copper	Chlorine	Copper chloride
Iron	Sulfur	Iron sulfide
Mercury	Oxygen	Mercury oxide
Silver	Iodine	Silver iodide

Table 2.10: Examples of Ionic compounds

Science as a human endeavour:

Water Fluoridation

The mineral fluoride occurs naturally on earth and is released from rocks into the soil, water, and air. Over the course of the 20th century, much research has positively correlated fluoride with the prevention of tooth decay. All water contains some fluoride though at concentrations not sufficient to provide safe guard against tooth decay. Water fluoridation has been introduced in 24 countries worldwide, including Australia as a public health strategy to reduce the prevalence of tooth decay.

The fluoride ionic compounds used to fluoridate water are derived from a mineral rock called fluorapatite ($Ca_5(PO_4)_3F$). There are three fluoride-releasing ionic compounds recommended in the Australian Drinking Water Guidelines for use in fluoridating water. These recommended fluoridating compounds are sodium fluoride (NaF), sodium fluorosilicate (Na_2SiF_6) and fluorosilicic acid (H_2SiF_6).

Questions

8. Many everyday substances are ionic compounds.

(a) The table below outlines combinations of cations and anions which react to form common ionic substances. Use this information to name the ionic compound.

Everyday Substance	Cation	Anion	Ionic Compound
Table Salt	Sodium	Chlorine	
Bath Salt	Magnesium	Chlorine	
Rust	Iron	Oxygen	
Toothpaste ingredient	Sodium	Fluorine	
Rock Salt	Calcium	Chlorine	

(5 marks) KA4

(b) Describe the formation of the ionic bond between the cation and anion in rock salt.

(3 marks) **KA1**

(c) Rock salt which has the formula $CaCl_{2,}$ compared to table salt which has the formula NaCl. Explain the difference in the number of chlorine atoms required.

(3 marks) **KA1**

(d) Draw a diagram to depict the electrostatic attraction between the cation and anion present in the common toothpaste ingredient.

(3 marks) **KA4**

Science understanding

Ionic compounds are continuous and are represented by empirical formulae.

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The geometrical structure of the ionic lattice is formed from the repeating electrostatic attraction between positive cations and negative anions. This results in the formation of solid ionic **salts**.

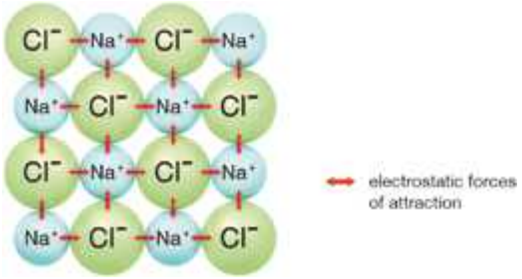


Figure 2.28: 2D model for ionic bonding

The ions are arranged so that the cation to anion attraction is maximised and the anion to anion, or cation to cation repulsion is minimised.

Figure 2.29: Ionic bonds between ions

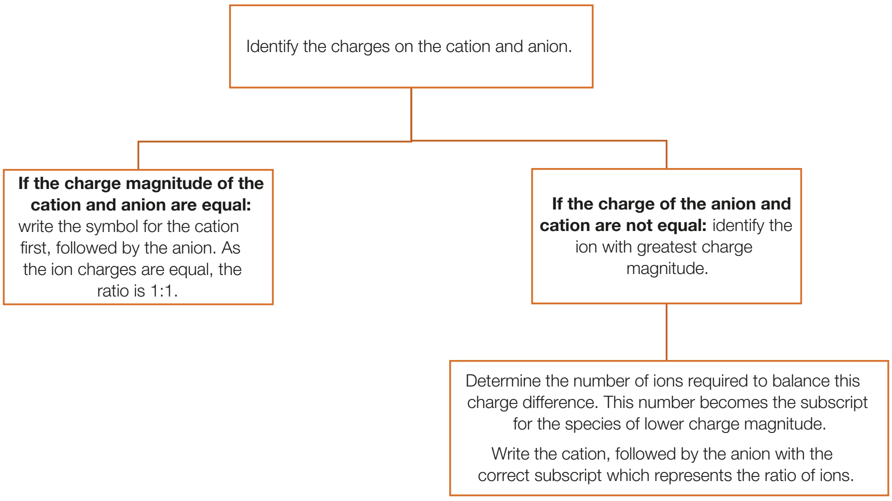
Anions and cations ionically bond in a ratio to allow the formation of electrically neutral compounds. The total number of negative charges, must equal the total number of positive charges. To form sodium chloride, only one chloride anion is needed for each sodium cation. In contrast, two chloride anions are needed for each magnesium cation to account for the difference in charge magnitude.

$$Na^+ + Cl^- \rightarrow NaCl$$

$$Mg^{2+} + Cl^- \rightarrow MgCl_2$$

This principle is the basis for understanding the process of writing empirical formulae for ionic compounds. Empirical formulae are the simplest ratio of elements in a compound, as ionic compounds do not form discrete units like molecules. Specifically, for an ionic compound it communicates the simplest ratio of ions. The subscript indicates the number of atoms needed of each ion to balance the charge.

The following procedure is followed to determine the empirical formulae for ionic compounds.



Identify the charges on the cation and anion.

If the charge magnitude of the cation and anion are equal: write the symbol for the cation first, followed by the anion. As the ion charges are equal, the ratio is 1:1.

If the charge of the anion and cation are not equal: identify the ion with greatest charge magnitude.

Determine the number of ions required to balance this charge difference. This number becomes the subscript for the species of lower charge magnitude.

Write the cation, followed by the anion with the correct subscript which represents the ratio of ions.

Example

1:1 Ratio - Potassium iodide

Identify the charges on the cation and anion.	K^+, I^-
When charges are equal, write the formula write the cation followed by the anion.	KI

The empirical formula for potassium iodide is KI.

1:2 Ratio – Lithium sulfide

Identify the charges on the cation and anion.	Li^+, S^{2-}
When the charges are not equal, identify the ion with the greatest charge.	S^{2-}
Determine the number of ions required to balance this charge difference.	2 × Li^+
This becomes the subscript for this species.	Li_2
Write the cation, followed by the anion with the correct subscript.	Li_2S

The empirical formula for lithium sulphide is Li_2S.

1:3 Ratio – Aluminium chloride

Identify the charges on the cation and anion.	Al^{3+}, Cl^-
When the charges are not equal, identify the ion with the greatest charge.	Al^{3+}
Determine the number of ions required to balance this charge difference.	3 × Cl^-
This becomes the subscript for this species.	Cl_3
Write the cation, followed by the anion with the correct subscript.	$AlCl_3$

The empirical formula for aluminium chloride is $AlCl_3$

2:3 Ratio – Aluminium oxide

Identify the charges on the cation and anion.	Al^{3+}, O^{2-}
When the charges are not equal, identify the ion with the greatest charge.	Al^{3+}
Determine the lowest common multiple (LCM) for both ion charges. Multiply the charge for each species by the value required to equal the LCM.	LCM is 6 2 × Al^{3+} 3 × O^{2-}
This becomes the subscript for this species.	$Al_2 : O_3$
Write the cation, followed by the anion with the correct subscript.	Al_2O_3

The empirical formula for aluminium oxide is Al_2O_3.

Oxidation States – Transition Metals

The total number of electrons that an atom either gains or releases in order to form a chemical bond with another atom is represented using the atom's **oxidation state**. Oxidation states are further explained in topic 6. Many **transition metals** can possess multiple oxidation states. As the transition metals have several electrons with similar ionisation energies, different conditions will result in the release of one or more electrons. For these elements, the oxidation state or number of electrons released is represented by Roman numerals placed within a bracket.

Example

1:1 Ratio – Iron(II) oxide

Identify the charges on the cation and anion.	Fe^{2+}, O^{2-}
When the charges are not equal, write the cation, followed by the anion with the correct subscript.	FeO

The empirical formula for iron(II) oxide is FeO.

2:3 Ratio – Iron(III) oxide

Identify the charges on the cation and anion.	Fe^{3+}, O^{2-}
When the charges are not equal, identify the ion with the greatest charge.	Fe^{3+}
Determine the lowest common multiple (LCM) for both ion charges. Multiply the charge for each species by the value required to equal the LCM.	LCM is 6 $2 \times Fe^{3+}$ $3 \times O^{2-}$
This becomes the subscript for this species.	$Fe_2 : O_3$
Write the cation, followed by the anion with the correct subscript.	Fe_2O_3

The empirical formula for iron(III) oxide is $Fe_2O_{3.}$

Polyatomic ions

All the above ionic compounds are comprised of **monoatomic** ions. Monatomic ions are ions derived from one atom. **Polyatomic** ions are those composed of more than one atom. Polyatomic ions arise when a covalent molecule undergoes an electron exchange. Most polyatomic ions are negative resulting from the gain of electron(s), with the ammonium (NH_4^+) and hydronium (H_3O^+) cations as notable exceptions.

Polyatomic Ion	**Formula unit**
hydroxide anion	OH^-
nitrate anion	NO_3^-
nitrite anion	NO_2^-
sulfate anion	SO_4^{2-}
sulfite anion	SO_3^{2-}
carbonate anion	CO_3^{2-}
bicarbonate anion	HCO_3^-
phosphate anion	PO_4^{3-}
acetate anion	CH_3COO^-
ammonium cation	NH_4^+
hydronium cation	H_3O^+

Table 2.11: Common Polyatomic ions and their formulae

Polyatomic ions behave as distinct units with an overall charge when determining their interaction with a positive metal cation. To determine the formulae of an ionic compound containing a polyatomic ion the same process is followed that is outlined above for monoatomic ions. A notable difference is when a subscript is applied to the polyatomic ion in the ionic formula, the ion is enclosed with parenthesis "()".

Example

1:1 Ratio – Copper(II) carbonate

Identify the metal cation and its charge.	Cu^{2+}
Identify the polyatomic ion and its charge.	SO_4^{2-}
When charges are equal, write the formula write the cation followed by the anion.	$CuSO_4$

The empirical formula for copper sulfate is $CuSO_4$.

1:2 Ratio – Magnesium nitrate

Identify the metal cation and its charge.	Mg^{2+}
Identify the polyatomic ion and its charge.	NO_3^-
When the charges are not equal, identify the ion with the greatest charge.	Mg^{2+}
Determine the number of ions required to balance this charge difference.	$2 \times NO_3^-$
This becomes the subscript for this species. Remember the use of brackets if the subscript is applied to the polyatomic ion.	$Mg : (NO_3)_2$
Write the cation, followed by the anion with the correct subscript.	$Mg(NO_3)_2$

The empirical formula for magnesium nitrate is $Mg(NO_3)_2$,

2:3 Ratio – Aluminium sulfate

Identify the metal cation and its charge.	Al^{3+}
Identify the polyatomic ion and its charge.	SO_4^{2-}
When the charges are not equal, identify the ion with the greatest charge.	Al^{3+}
Determine the lowest common multiple (LCM) for both ion charges. Multiply the charge for each species by the value required to equal the LCM.	LCM is 6 $2 \times Al^{3+}$ $3 \times SO_4^{2-}$
This becomes the subscript for this species. Remember the use of brackets if the subscript is applied to the polyatomic ion.	$Al_2 : (SO_4)_3$
Write the cation, followed by the anion with the correct subscript.	$Al_2(SO_4)_3$

The empirical formula for aluminium sulfate is $Al_2(SO_4)_3$.

Questions

9. Metals commonly release electrons and combine with non-metals (which accept the valence electrons) to result in the formation of electrically neutral ionic compounds. The ratio of cation : anion required is used to determine the empirical formula of the compound.

(a) Use this understanding to write the empirical formula of the following compounds:

Ionic Compound	**Empirical Formula**
sodium phosphide	
aluminium sulphide	
beryllium chloride	
gallium nitride	
calcium oxide	
lithium bromide	
magnesium fluoride	
potassium phosphide	

(8 marks) **KA4**

(b) Many transition metals can possess multiple oxidation states. For these elements, the oxidation state or number of electrons lost is represented by Roman numerals placed within a bracket. Use this understanding to write the empirical formula of the following compounds:

Ionic Compound	**Empirical Formula**
lead(II) sulfide	
copper(I) arsenide	
iron(III) oxide	
iron(II) bromide	
mercury(II) oxide	
chromium(III) chloride	
tin(IV) chloride	
iron(II) sulfide	

(8 marks) **KA4**

(c) Many ionic compounds contain polyatomic ions. Write the formula for the following polyatomic ionic compounds:

Ionic Compound	**Empirical Formula**
magnesium hydroxide	
sodium nitrate	
vanadium(V) phosphate	
potassium sulfate	
calcium nitrite	
sodium bicarbonate	
copper(I) carbonate	
lead(II) sulfite	
beryllium acetate	
ammonium hydroxide	

(10 marks) **KA4**

Science understanding

The properties of ionic compounds can be explained using the model for ionic bonding.

Hard and brittle solids

Ionic compounds form a continuous 3D lattice from the electrostatic attraction between cations and anions. This most commonly results in salts that are solids at room temperature.

Science as a human endeavour:

Ionic Liquids

An ionic liquid is a salt in the liquid state. While liquids such as water are comprised of electrically neutral molecules, ionic liquids are made predominantly of ions.

The biological activity of ionic liquids and their subsequent application in pharmaceuticals and medicine has gained much momentum in the scientific community. They have been described as the basis of a quiet revolution in material science.

There has been much documented development in this research area including the investigation of ionic liquids as antibacterial and antifungal agents. During a time where there are strains of pathogens that are resistant to many existing microbial agents the influence of this research may be significant for society.

The orderly arrangement of ions in the lattice gives rise to the various shapes of the naturally occurring ionic **crystals**.

Science as a human endeavour:

Transition Metals

Transition metals form the *d* block of the periodic table. They are versatile materials with applications in diverse fields from medicine to construction and energy due to their desirable properties and multiple oxidation states. Ionic compounds including ions of **transition** metals often result in a brilliance of colour. The colours arise when electrons within the metals in the compounds absorb light of a specific wavelength and move to higher energy levels. Examples include azurite, a copper crystal, and cinnabar, a mercury crystal as shown below in Figure 2.30.

Figure 2.30: Azurite – a copper crystal: $Cu_3(CO_3)_2(OH)_2$ and Cinnabar – a mercury crystal: HgS

Ionic compounds containing transition metals are extracted and isolated for extensive use in pigments in paints and to colour glass, ceramics and enamel.

Within the lattice the opposing charges attract one another, forming strong ionic bonds. Likewise, the like charges repel each other, maintaining the regular lattice. The result is a highly stable structure. The stability of the ionic lattice gives them a high degree of hardness.

However, unlike metallic structures they are **brittle** compounds. Brittle materials fracture due to their low compressive strength when subjected to mechanical stress. They have minimal ability to undergo elastic deformation. When ionic compounds are exposed to compressive or mechanical stress, the layers of ions shift. This forces the ions of like charges close to one another. The induced repulsion causes the solid to shatter across a smooth plane of the material.

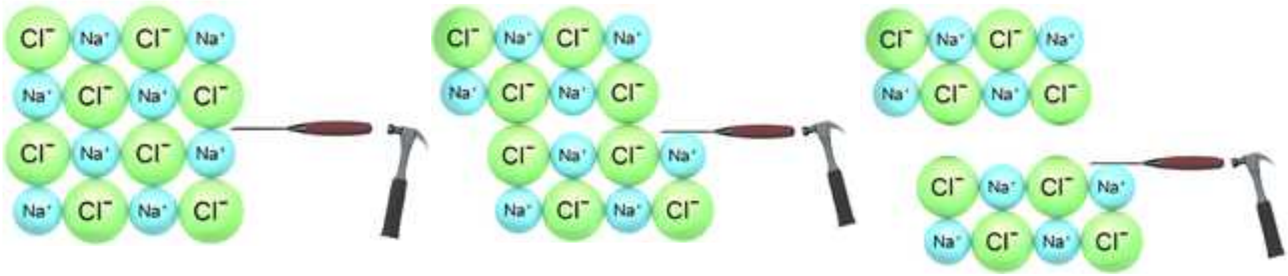


Figure 2.31: Change in position of ions when mechanical stress is applied

Melting and boiling point

The melting point of an ionic solid is dependent on the strength of the bonding present. For ionic solids it is the strong electrostatic forces of attraction, the ionic bonds between positive cations and negative anions, that hold the ionic lattice together. Thus, ionic compounds are materials with high melting points. When ionic bonds are exposed to sufficient heat energy the cations and anions will separate forming a molten ionic compound.

The strength of ionic attraction is increased when the ions contained in the lattice are of greater charge magnitude. This results in a greater requirement of applied energy to separate them, increasing the compound's melting point. If heated to their melting point, the electrostatic attraction will be overcome resulting in the separation of cations and anions into a liquid phase. The liquid formed is referred to as a molten ionic compound.

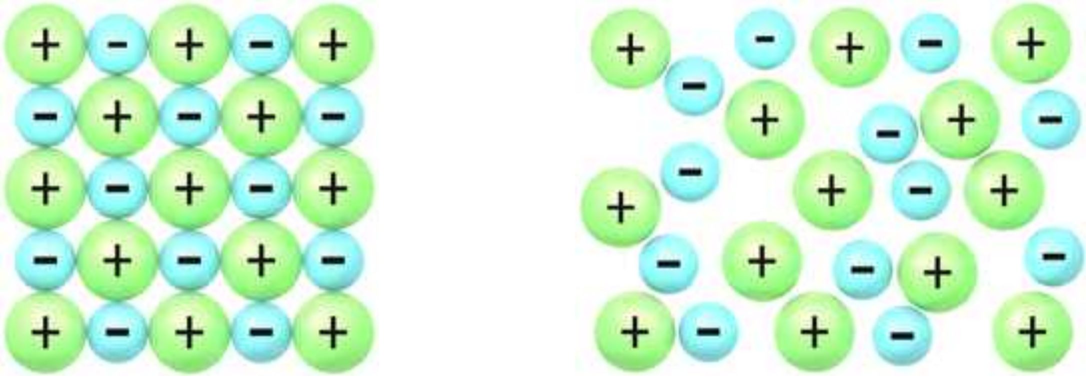


Figure 2.32: Ionic compounds in the solid (left) and molten or liquid (right) phase

Once the ionic lattice is molten and heat is continually applied, it eventually reaches a temperature at which the vapour pressure is great enough to form bubbles inside the body of the liquid. When this boiling point of an ionic compound is reached the anions and cations separate further forming a gas.

Electrical conductivity

Electrical conductivity is the measure of the ease of which an electric charge can pass through a material. Solid ionic compounds do not conduct electricity as the ions are in fixed positions and therefore cannot carry charge. All ionic compounds are **electrolytes**, compounds which produce ions when dissolved into an aqueous solution with water (if soluble) or when molten. ***Solubility*** *of compounds is explored in Topic 4.* Thus, electrolytes contain ions that are free to move, ions that can carry charge. Molten ionic compounds and aqueous ionic solutions are conductors of electricity.

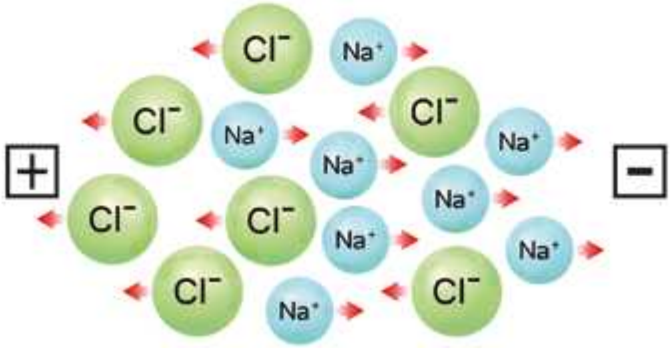


Figure 2.33: Electrical conductivity in a molten ionic electrolyte

Within the electrolyte, the ion's charge magnitude or the number of ions present will impact the magnitude of current they can conduct.

Example

Comparing conductivities

Magnesium sulfate and sodium chloride are both soluble ionic compounds. When placed in water, they dissolve to release ions that can carry charge. An electrolyte solution of magnesium sulfate contains Mg^{2+} and SO_4^{2-} ions, compared to a sodium chloride electrolyte solution which produces Na^+ and Cl^- ions. The Mg^{2+} and SO_4^{2-} ions are doubly charged in comparison to Na^+ and Cl^- ions. Consequently, a magnesium sulfate electrolyte can carry two times the **amperes** of current compared to a sodium chloride electrolyte.

Sodium sulfate, another soluble ionic compound, releases Na^+ and SO_4^{2-} ions when dissolved in water in a 2:1 based on it's empirical formula, Na_2SO_4. The resulting electrolyte has greater conductivity in comparison to NaCl at the same concentration for two reasons. Sodium sulfate has twice as many Na^+ ions to carry charge, and secondly that SO_4^{2-}ions have a greater charge magnitude compared to the Cl^- ions present in sodium chloride.

Questions

10. Potassium iodide blocks uptake by the thyroid gland of radioactive iodine that may be released during a nuclear radiation disaster.

    (a) Potassium iodide is a ionic solid. Describe how a solid results from the interaction of cations and anions.

....

....

....

....

.... *(3 marks)* **KA1**

(b) Draw an electron transfer diagram to depict the formation of the ionic bond in potassium iodide.

(3 marks) **KA4**

(c) Write the formula of potassium iodide.

.... *(1 mark)* **KA4**

(d) Explain the relatively high melting point (671°C) of potassium iodide.

(2 marks) **KA1**

(e) Potassium iodide is highly soluble. Predict the conductivity of this solution as compared to other ionic compounds, justifying your response.

(2 marks) **KA1**

(f) Potassium iodide has low compressive strength. Explain what is meant by this statement.

(2 marks) **KA1**

11. Some reactive metals are extracted from their ores using electrolysis. A source of direct current is connected to a pair of inert electrodes immersed in a molten ionic compound.

(a) Explain what happens to the ionic lattice when it becomes molten.

(2 marks) **KA1**

(b) Molten sodium chloride readily conducts electricity, allowing it to be used as an electrolyte. Compare and explain the difference in conductivity of solid and molten sodium chloride.

(3 marks) **KA1**

Science understanding

Non-metallic atoms share electrons to form covalent bonds.

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Main-group elements usually adhere to the octet rule, exhibiting a preference to complete their valence shell with eight electrons. This stable arrangement is either comprised from the atom's own electrons or electrons that are released or shared with another atom. It requires eight electrons to fill the *s* and *p* orbitals of the corresponding energy level, to attain their noble gas configuration.

Unlike ionic bonding, covalent bonding does not involve the exchange of valence electrons to gain octet stability. Covalent bonding occurs when valence electron pairs are shared by non-metallic atoms. Non-metallic elements have a preference to share electrons due to their high electronegativities. In these instances strong electrostatic attraction arises between the bonding pair of valence electrons and the positively charged nuclei of both participating atoms. This attraction is termed a covalent bond.

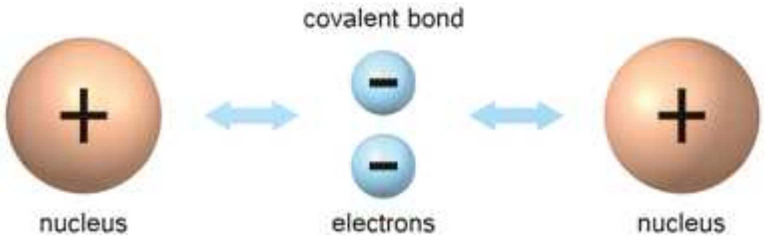


Figure 2.34: Covalent bond

Single covalent bonds

A single covalent bond involves two atoms sharing one pair of electrons to complete their octet.

Group VII elements have incomplete valence shells with 7 electrons. By combining with an atom of another group VII element a bonding pair of electrons is formed and both atoms complete their valence shell and attain octet stability. The bonding electrons are electrostatically attracted to the positive nuclei of both atoms, covalently bonding the atoms together. Two chlorine atoms can complete their 3*p* orbitals by sharing a bonding pair of electrons, forming a covalent bond. Each chlorine atom also has three non-bonding pairs of electrons.

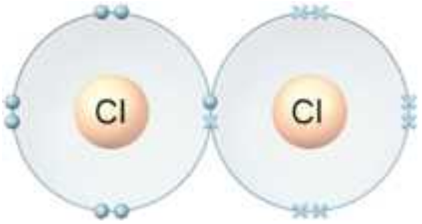


Figure 2.35: a single covalent bond between two chlorine atoms

Double covalent bonds

Atoms can share more than one pair of electrons. A double covalent bond involves two atoms sharing two pairs of electrons to complete their octet.

Group VI elements have incomplete valence shells with 6 electrons. By combining with an atom of another Group VI element, two bonding pairs of electrons are formed, and both atoms complete their valence shell and 3*p* orbital occupancy. Two atoms of oxygen can share four electrons (two from each atom), with the electrostatic attraction between the two bonding electron pairs and positive nuclei of the atoms resulting in the formation of a double bond. Both oxygen atoms possess two non-bonding pairs of electrons.

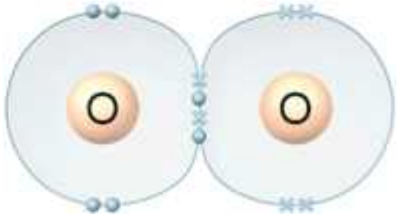


Figure 2.36: a double covalent bond between two oxygen atoms

Triple covalent bonds

Group V elements have incomplete valence shells with 5 electrons. When atoms of these elements combine, three bonding pairs of electrons are shared, six electrons in total. Two nitrogen atoms can bond in this manner, resulting in the following arrangement of electrons. Each nitrogen atom also has one non-bonding pair of electrons.

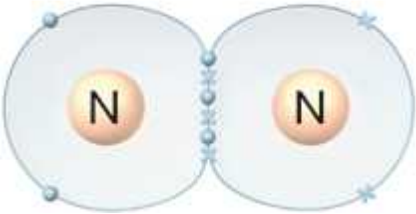


Figure 2.37: a triple covalent bond between two nitrogen atoms

Bond Strength

As the number of bonding pairs of electrons increases, the strength of the intramolecular attraction also increases. A single covalent bond is the weakest as the interaction only involves two electrons. A double bond has intermediate strength based on the electrostatic attraction between two bonding pairs of electrons. The triple covalent bond is the strongest due to the interaction of three bonding pairs of energy. As the bond strength increases, the quantity of energy required to break the bond also increases.

Science as a human endeavour:

Haber Process

As a triple covalently bonded diatomic molecule, the intramolecular forces that hold nitrogen together are incredibly strong. A significant amount of energy is required to overcome these bonds, in nature sufficient energy is provided by lightening or volcanoes. An internal combustion engine can also break this bond, allowing nitrogen to combine with oxygen in the formation of other nitrogen containing molecules.

Nitrogen is a crucial element required for protein synthesis, however plants cannot absorb it in its molecular form. Conversion of nitrogen into useable forms, called ‘nitrogen fixation’, is therefore essential for plant growth. Nitrogen based fertilisers to support plant growth often contain another nitrogen based molecule, ammonia. Ammonia is commercially produced via the Haber Process. The Haber process involves the use of high temperatures and pressure to convert its constituent elements, hydrogen and nitrogen into ammonia.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Named after Fritz Haber, a gifted chemist of the early 20th century the Haber process supported the growing need of nitrogenous fertiliser to assist with food production for rapid population growth experienced during this time.

Covalent Compounds: Molecules

Covalent bonds are the intramolecular attractions that result in the formation of covalent compounds. Covalent compounds are often described as molecules or molecular compounds. Molecules are discrete units that interact with adjacent molecules via intermolecular attractive forces. They can be comprised of two or more atoms, of the same or differing elements. Molecules comprised of the same element are **homonuclear**. Those comprised of differing elements are termed **heteronuclear** molecules.

Chlorine, oxygen and nitrogen molecules are all examples of **diatomic** molecules. Diatomic molecules are composed of only two atoms from the same or different elements. Seven elements exist as homonuclear diatomic gaseous molecules at room temperature, H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2.

To represent the electrons that are involved in a molecules covalent bonding an **electron dot diagram** is drawn. These diagrams are often referred to as Lewis dot diagrams, named after their creator Gilbert. N. Lewis. In 1916, Lewis first proposed a chemical bond involving the sharing of electrons. He described the “cubical atom”, as a cube has 8 corners each representing the valence electrons of an atom that can be used to form a bond. Today we know that electron orbitals do not take a cubic shape, but Lewis electron dot diagrams remain a useful tool.

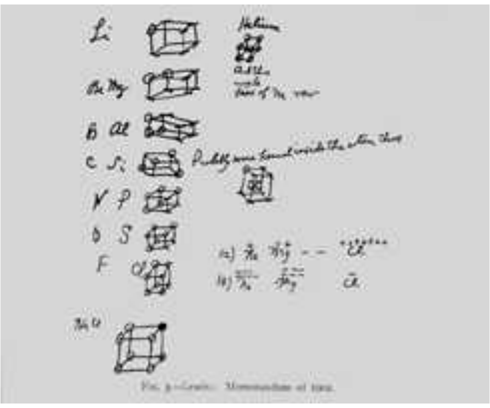


Figure 2.38: Lewis’s graphical representations of the role of electrons in atomic structures (Science History Institute)

Lewis introduced simple symbols to denote an atom’s valence electrons. These symbols are known as electron dot diagrams and can be used as the basis of a Lewis structure.

Electron dot diagrams

Electron dot diagrams are shorthand representations of the valence electrons of an atom with the number of dots equalling the number of valence electrons. Atoms typically form a characteristic number of covalent bonds in compounds, as summarised in Figure 2.39.

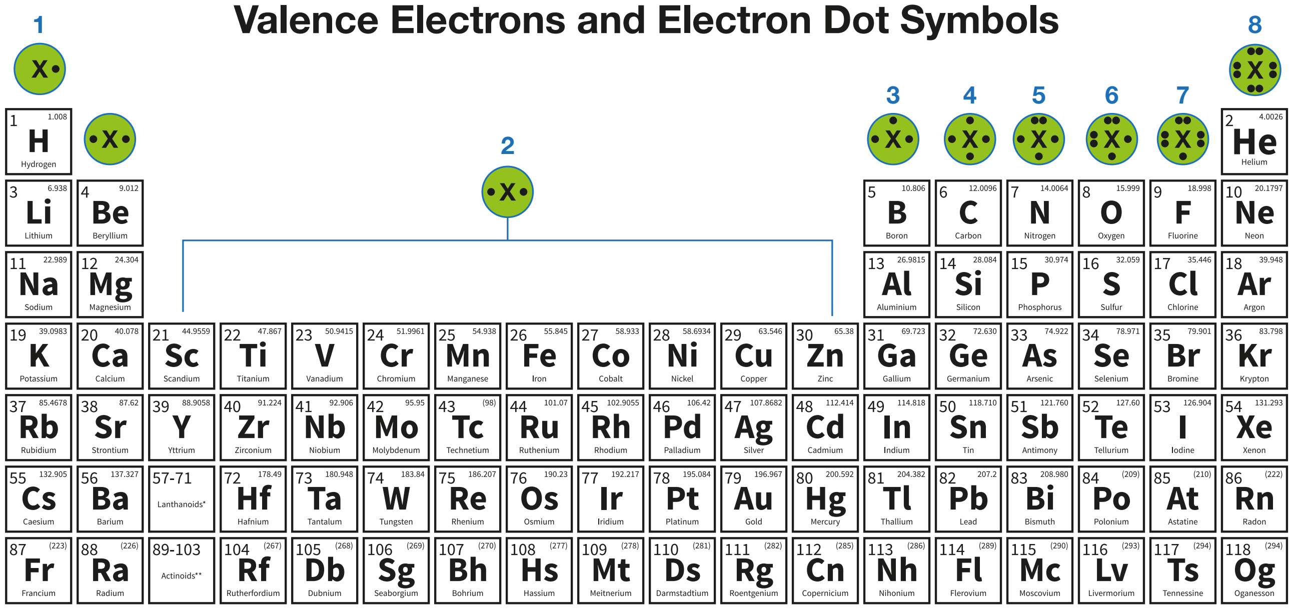


Figure 2.39: Valence electrons and electron dot symbols

An element's electron configuration is used to determine the number of valence electrons, from which a dot diagram can then be drawn. Common examples of electron dot diagrams for monoatomic non-metallic elements are shown below in Table 2.12. Note that electrons exist in pairs, but only as the 5th–8th electrons are added.

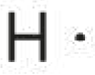
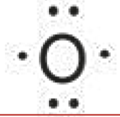
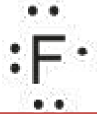
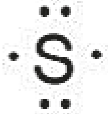
Element	Electron configuration	Number of valence electrons	Electron Dot Diagram
Hydrogen	$1s^1$	1	
Nitrogen	$1s^22s^22p^3$	5	
Oxygen	$1s^22s^22p^4$	6	
Fluorine	$1s^22s^22p^5$	7	
Phosphorous	$1s^22s^22p^63s^23p^3$	5	
Sulfur	$1s^22s^22p^63s^23p^4$	6	
Chlorine	$1s^22s^22p^63s^23p^5$	7	

Table 2.12: Electron dot diagrams for common non-metallic elements

Electron dot diagrams can also be used to show how atoms covalently bond by sharing electrons. Homonuclear diatomic molecules can be represented in this way. A single, double or triple covalent bond between the atoms is denoted by the number of bonding electron pairs.

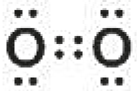
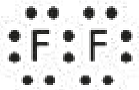
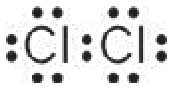
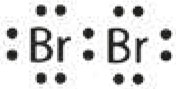
Molecule	Formula Unit	Number of bonding electron pairs	Electron Dot Diagram
Hydrogen	H_2	1	
Nitrogen	N_2	3	
Oxygen	O_2	2	
Fluorine	F_2	1	
Chlorine	Cl_2	1	
Bromine	Br_2	1	

Table 2.13: Electron dot diagrams for common homonuclear molecules

Electronegativity

The central atom of a molecule is the atom with highest **covalence** and is usually the element of lowest electronegativity. Covalence refers to the number of pairs of electrons that an atom can share with neighbouring atoms. Hydrogen and group VII atoms are rarely the central atom due to this. Electronegativity is a chemical property that describes the tendency of an atom to attract bonding electrons towards itself. Electronegativity is a measure of the electrostatic attraction between the atom's positive nucleus and a negative bonding electron.

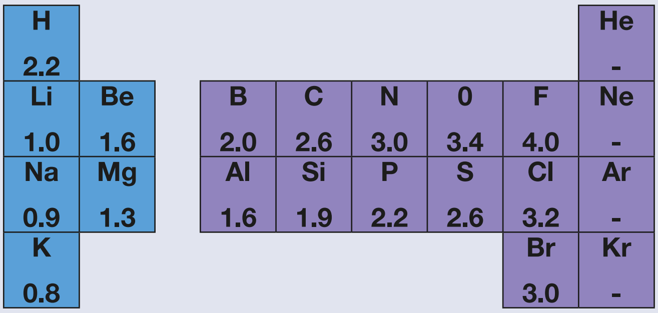
Science as a human endeavour:

Linus Pauling and Electronegativity

Electronegativity was first proposed by Linus Pauling in 1932 following the development of valence bond theory. Pauling was an American theoretical physical chemist who used quantum mechanics to understand and describe chemical bonding. Pauling remains the only individual to be awarded two unshared Nobel prizes.

Electronegativity cannot be directly measured and thus must be calculated as a scaled value from other atomic properties. The most widely accepted calculation was also proposed by Pauling producing a relative set of atomic values from 0.7 (Francium) to 4.0 (Fluorine). The greater the number, the greater the element's electronegativity.

The "Pauling scale" allows the trend of electronegativity to be explored quantitatively.



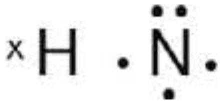
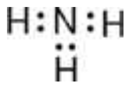
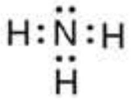
H **2.2**							**He** **-**
Li **1.0**	**Be** **1.6**	**B** **2.0**	**C** **2.6**	**N** **3.0**	**0** **3.4**	**F** **4.0**	**Ne** **-**
Na **0.9**	**Mg** **1.3**	**Al** **1.6**	**Si** **1.9**	**P** **2.2**	**S** **2.6**	**Cl** **3.2**	**Ar** **-**
K **0.8**						**Br** **3.0**	**Kr** **-**

Figure 2.40: Pauling values (electronegativity) of the main group elements

For heteronuclear molecules a sequential process is followed to draw an electron dot diagram. The steps for this are outlined below.

1. Write the chemical symbol for the molecule.
2. Determine the number of valence electrons for each atom in the molecule.
3. Determine the number of electrons each atom needs to attain its octet configuration.
4. Determine the number of bonds needed in the molecules. Step 2 identifies the number of electrons needed, and step 3 identifies the number of electrons required to complete the octet of each atom.
5. Identify the central atom, which is the atom that has the highest **covalence**.
6. Complete the octets around each of the outside atoms.
7. Place remaining electrons around the central atom. These are the **non-bonding** or **lone electrons** which play an important role in the subsequent molecule shape. *This is explained in Topic 3.*
8. Check that the total number of electrons in the diagram matches the total number of valence electrons that the constituent atoms started with.

Example

Ammonia	
1. Write the chemical symbol for the molecule.	NH_3
2. Determine the number of valence electrons for each atom in the molecule.	
3. Determine the number of electrons each atom needs to attain its octet configuration.	Nitrogen requires 3 electrons to complete it's octet. Three hydrogens are present, allowing each to share one electron to form a bonding pair.
4. Determine the number of bonds needed in the molecules. Step 2 identifies the number of electrons needed, and step 3 identifies the number of electrons required to complete the octet of each atom.	Three hydrogens are present, allowing the formation of 3 bonds.
5. Identify the central atom. It is usually the atom that has the highest **covalence**.	Nitrogen is the central atom as it has the highest covalence.
6. Complete the octets around each of the outside atoms.	
7. Place remaining electrons around the central atom. These are the non-bonding or lone electrons which play an important role in the subsequent molecule shape. *This is explained in Topic 3.*	

Lewis structure diagrams

An electron dot diagram is used as the basis for a Lewis structure diagram which transforms the bonding electron pair to be represented as one or more straight lines depicting the single, double or covalent bond. Lewis structure diagram for common molecules are shown in Table 2. Importantly non-bonding electron pairs are still included.

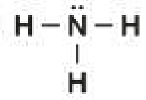
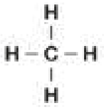
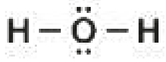
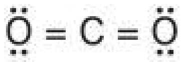
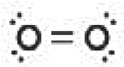
Molecule	Molecular formula	Lewis Structure Diagram
Ammonia	NH_3	
Methane	CH_4	
Water	H_2O	
Carbon Dioxide	CO_2	
Oxygen	O_2	
Nitrogen	N_2	

Table 2.14: Lewis Structure diagrams for common non-metallic molecules

Electronegativity is demonstrative of periodicity. As a trend, electronegativity increases from left to right on the periodic table, and from the bottom to the top of a group (excluding Group VIII elements). The element with the greatest electronegativity is fluorine.

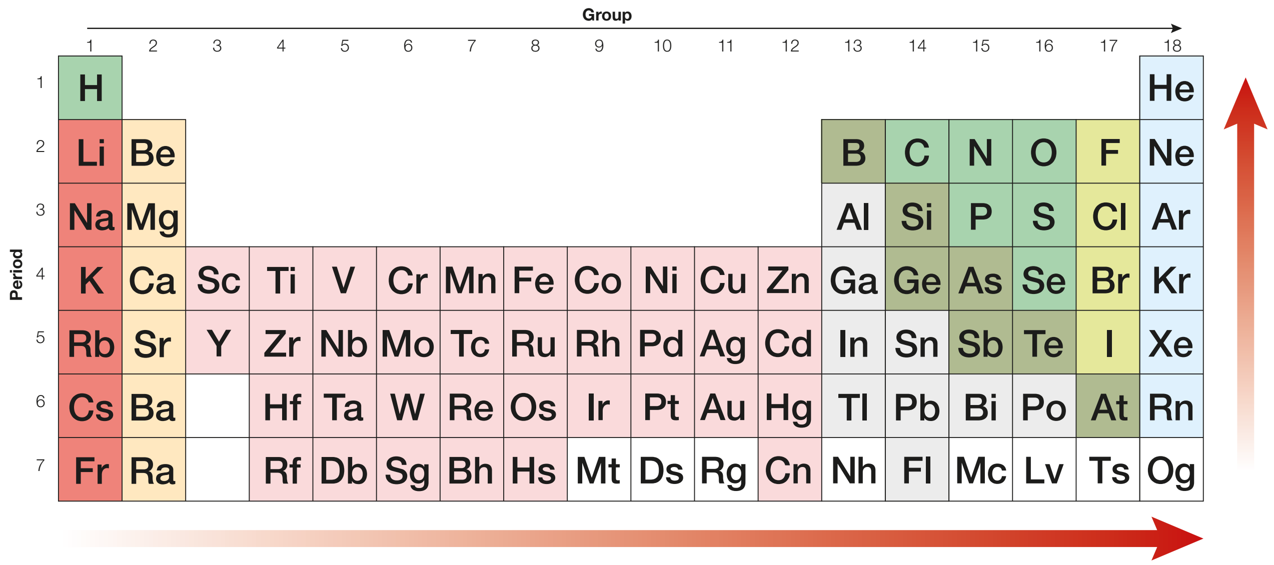


Figure 2.41: Periodicity of electronegativity

The magnitude of an atom's electronegativity correlates with a number of factors.

Variation along a period– Ionisation energy increases across a period. The higher the ionisation energy of an atom, the higher it's electronegativity (with the exception of the Group VIII elements) due to a high electron affinity.

Variation down a group – electronegativity decreases down a group due to the increase in atomic radius. The greater the size of the atom, the lesser the tendency of its nucleus to attract bonding electrons due to the greater distance between the opposing charges.

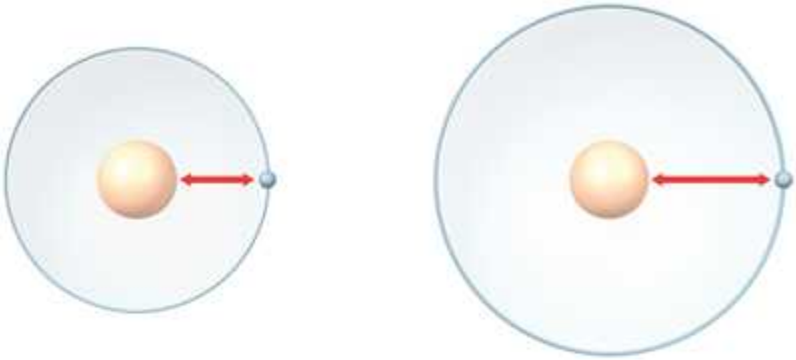


Figure 2.42: Atomic radii – smaller atomic radius (left) and larger atomic radius (right)

Atomic number – electronegativity decreases with an increase to an atom's atomic number. As the atomic number increases, there is a greater number of positive protons present to attract bonding electrons. However, the electrons close to the nucleus shield the nucleus from exhibiting this attraction. This is termed **electron shielding** and decreases the electronegativity of an atom.

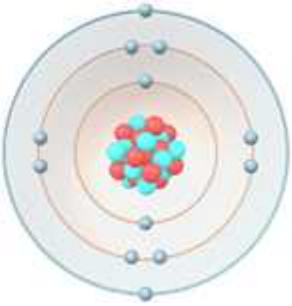


Figure 2.43: Electron shielding

Questions

12. Oxygen is vital for the sustainment of life on Earth, and is the second most abundant atmospheric gas, with air commonly containing approximately 21% oxygen.

    (a) Draw a Lewis dot diagram for an atom of oxygen.

    (2 marks) **KA4**

    (b) Explain the formation of the covalent bond between two atoms of oxygen.

    (2 marks) **KA1**

    (c) Draw a Lewis dot diagram for a oxygen molecule showing all non-bonding electron pairs.

    (2 marks) **KA4**

(d) Oxygen reacts with nitrogen in the high temperatures provided by an internal combustion engine.

$$N_2 + O_2 \rightarrow 2NO$$

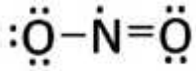
Nitric oxide can react further in the lower atmosphere with oxygen to produce nitrogen dioxide.

$$NO + O_2 \rightarrow NO_2$$

(i) Balance the equation above.

(1 mark) **KA4**

(ii) Nitrogen dioxide is an exception to the octet rule as the total number of valence electrons is an odd number. Thus it's valence structure is as below:



Compare the two types of covalent bonds present in the nitrogen dioxide molecule.

....

....

.... *(2 marks)* **KA1**

(iii) Nitrogen is less electronegative than oxygen and therefore is the central atom. Explain what is meant by the term electronegative.

....

....

.... *(1 mark)* **KA1**

(iv) Explain why nitrogen is less electronegative than oxygen.

....

....

.... *(2 marks)* **KA1**

(e) Other gases exist in the atmosphere as covalent molecules. For each molecule complete its Lewis structure diagram.

Name	Chemical formula	Lewis structure diagram
Carbon dioxide	CO_2	
Methane	CH_4	
Hydrogen	H_2	
Oxygen	O_2	

(8 marks) **KA4**

Science understanding

A covalent bond may be polar or non-polar.

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A non-metallic atom's electronegativity will impact the way electrons are shared with another non-metallic atom in the formation of a covalent bond. The distribution of electrons in the bond is dependant on the atoms that are present. There may be partial transfer of **electron density** from one atom to another. This occurs when the bonding atoms have differing electronegativities, the electrons are not shared equally, and partial charges arise. This is termed **bond polarity**, thus allowing covalent bonds to be described as either non-polar or polar in nature.

Non-polar covalent bonds

Non-polar covalent bonds result from the combination of two non-metallic atoms of the same electronegativity sharing electrons. The bonding electrons have equal attraction to both positive nuclei therefore are held an equal distance apart. No **partial charge** or **dipole** is introduced as the charges are equally distributed.

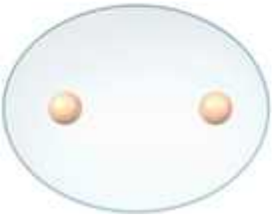

Figure 2.44: Non-polar molecule with equally distributed electron density.

Most commonly this occurs when atoms of the same element combine. However, in some instances the difference in two element's electronegativity is negligible, which therefore still results in a functionally non-polar covalent bond.

Example

Chlorine

In the formation of the homonuclear chlorine molecule (Cl_2), the bonding electrons have equal attraction to both positive nuclei therefore are held an equal distance apart. As no partial charge or dipole is introduced the chlorine molecule is non-polar.

Carbon – Hydrogen bond

Carbon has a high capacity for covalent bonding due to its outer shell valence of four. Carbon commonly bonds to hydrogen to give rise to hydrocarbon compounds, and many other organic compounds important for life on Earth. Carbon has a Pauling value of 2.5, a slight increase compared to hydrogen's value of 2.1. Due to the small magnitude of difference in the atom's electronegativity the C-H bond is deemed non-polar. *This is important when considering overall molecule polarity which will be explored in Topic 3.*

Polar covalent bonds

Covalent bonds that arise from the sharing of bonding electrons between atoms of differing electronegativity are termed **polar covalent bonds**. The difference in the atom's tendency to attract bonding electrons results in an unequal sharing of the electron pair. The result is a bond where the bonding electron pair is displaced towards the more electronegative atom.

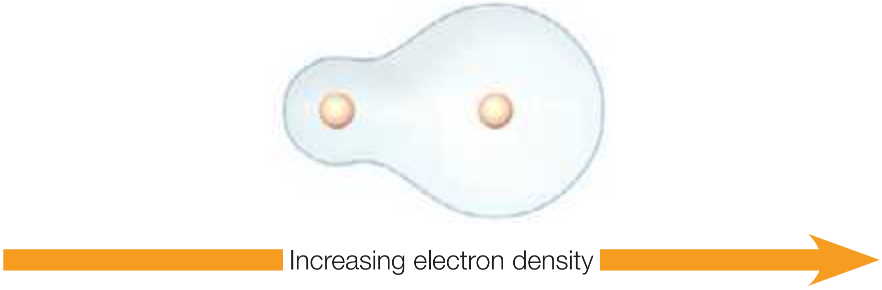

Figure 2.45: Induced dipole of a polar covalent bond

The more electronegative atom then obtains a partial-negative charge while the less electronegative atom has a partial-positive charge. The induced dipole moment is illustrated using negative and positive **delta** notation respectively. Delta notation comes from the Greek alphabet and is given the symbol δ, which appears as a superscript above the atomic symbol followed by a + or – dependant on the required charge. Delta notation is not required for non-polar molecules, as they do not contain partial charges. The greater the degree of electronegativity difference between the two bonding atoms, the greater the polarity of the bond, and the greater the partial charges acquired by each atom in the bond.

Partial charges have no fixed value. The symbol "δ+" could represent a very small partial positive charge, or a slightly larger one, but always less than a formal +1 charge.

Example

Hydrochloric acid (HCl)

In the formation of hydrochloric acid (HCl), the combination of hydrogen and chlorine involves two atoms of great electronegativity difference. Chlorine has a Pauling value of 3.0, significantly higher than hydrogen's 2.1. The bonding electron pair is displaced towards chlorine due to its greater tendency to attract the electron pair, resulting in an induced dipole moment.

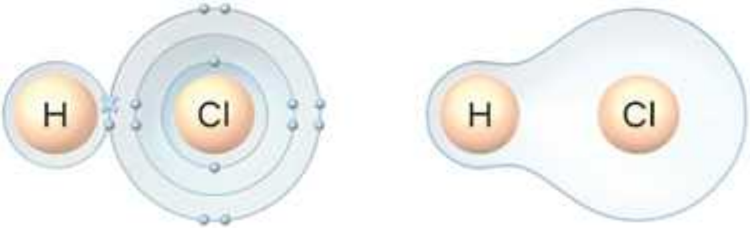

Figure 2.46: polarity of hydrogen chloride

The chlorine atom exhibits a partial negative charge as the bonding electron pair are closer to the chlorine atom. The hydrogen atom has a partial positive charge due to the greater distance between it and the negatively charged bonding electron pair.

$$\overset{\delta^+}{\mathrm{H}}\text{——}\overset{\delta^-}{\mathrm{Cl}}$$

As chemical bonds increase in polarity, they transform from pure covalent to ionic. The greater the polarity of the bond, the more ionic in character the bond is described. The atom of greater electronegativity attracts the bonding electrons forming the covalent bond, ionising the adjacent atom.

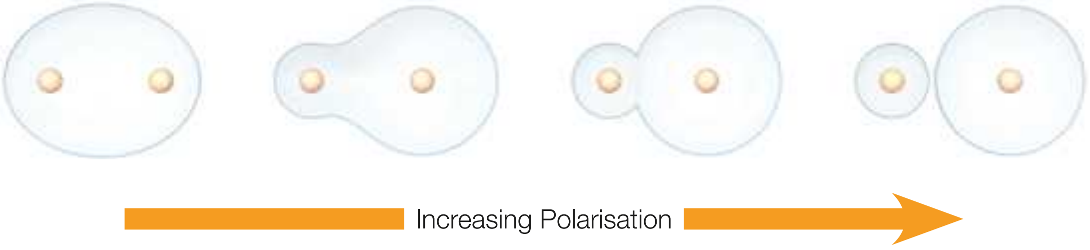

Figure 2.47: Changes in polarisation of a chemical bond

Questions

13. Hydrogen sulfide (H_2S) is found naturally in crude petroleum and natural gas. It is also produced through the bacterial breakdown of organic matter.

    (a) Draw a Lewis structure diagram for hydrogen sulfide.

    (2 marks) **KA4**

    (b) Indicate the polarity of the bonds in the molecule using delta notation (δ^+, δ^-).

    (1 mark) **KA4**

    (c) Explain why the hydrogen-sulfur bond is polar.

    (2 marks) **KA1**

    (d) Methane (CH_4) is another gas produced in the bacterial breakdown of matter.

    (i) Draw a Lewis dot diagram for methane.

    (2 marks) **KA4**

    (ii) State and explain the polarity of the bonds within the methane molecule.

    (2 marks) **KA1**

    (e) Hydrogen can be produced from methane via steam reforming. The process produces syngas (hydrogen and carbon monoxide) by reacting methane with water.

    (i) Draw a Lewis structure diagram for a hydrogen molecule.

    (2 marks) **KA4**

    (ii) Explain why all homonuclear diatomic molecules contain non-polar covalent bonds.

    (2 marks) **KA1**

Science understanding

Covalent bonding is found in molecular and non-molecular (continuous) substances.

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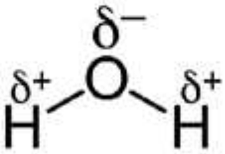
Covalent compounds: Molecules

As described earlier, covalent compounds are often described as molecules or molecular compounds. Molecules are discrete units, comprised of a fixed number of atoms. They also interact with adjacent molecules via intermolecular attractive forces.

Example

Water

Water is a molecular compound comprised of two hydrogen atoms covalently bonded to a central oxygen atom. The oxygen is the atom of greater electronegativity and higher valence allowing the hydrogen atoms to form a bonding electron pair with two of oxygen's valence electrons. The resulting bond is a polar covalent bond, due to the difference in electronegativity of hydrogen and oxygen. Such polarity induces a dipole, with partial negative charge on the oxygen atom, and partial positive charge on the hydrogen atoms.



The water molecules are discrete units, interacting with adjacent molecules to form intermolecular bonds. These intermolecular bonds give the liquid known as water many of its properties.

The physical properties of covalent compounds are explained in Topic 3.

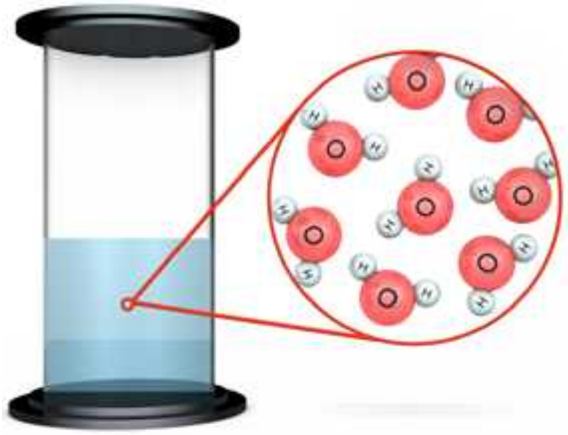


Figure 2.48: Water a molecular compound

Covalent compounds: continuous covalent networks

Covalent networks are large, rigid three-dimensional arrangements of atoms. The network lattice results from each atom forming covalent bonds with the nearest neighbouring atom at each valence electron site. The network of covalent bonds extends throughout the solid. Unlike molecules the number of atoms is not fixed and will depend of the size of the structure. The smallest amount of a network solid that can be identified as such is called a **formula unit**.

Carbon, silicon and boron all form covalent networks. Examples of covalent networks consisting of the same element are diamond and graphite, both of which are made entirely from carbon atoms. Covalent networks comprised of two elements include silica, boron carbide and silicon carbide.

Example

Carbon based covalent networks

Carbon is a non-metallic element with electron configuration $1s^2 2s^2 2p^4$, thus has a valence of four. This allows it to be the central atom in many configurations, giving rise to a wide variation of possible combinations with other atoms. In addition to combining covalently with other elements, carbon can also bond with itself, forming **allotropes**. Allotropes are different forms of the same element in the same state. Two of the most commonly known carbon allotropes are the covalent networks diamond and graphite.

Diamond

In diamond, each carbon shares electrons with four adjacent carbon atoms to form a covalent network. Covalent networks are similar in structure to ionic compounds forming a repeating three-dimensional structure as shown in the Figure 2.49 below.

Figure 2.49: Chemical structure of diamond

Graphite

Graphite is another carbon-cased covalent network. In graphite each carbon atom covalently bonds to three adjacent carbon atoms; the fourth is delocalised, and can move freely between the graphite layers. This bonding model forms layers with a hexagonal arrangement as shown in the Figure 2.50 below.

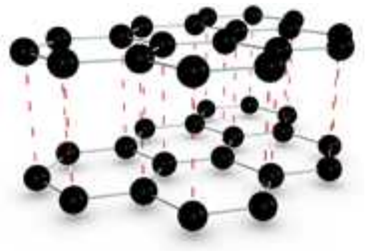


Figure 2.50: Chemical structure of graphite

Example

Silica

Silicon dioxide or silica is the main compound found in sand and arises from the covalent bonding between silicon and oxygen. Each silicon atom forms four covalent bonds with the adjacent oxygen atoms. There are two oxygen atoms to every silicon atom. This results in silica's formula unit of $(SiO_2)_n$ with *n* representing a large number of silicon and carbon atoms. Its bonding arrangement results in a similar structure to diamond with no free electrons and a repeating three-dimensional structure.

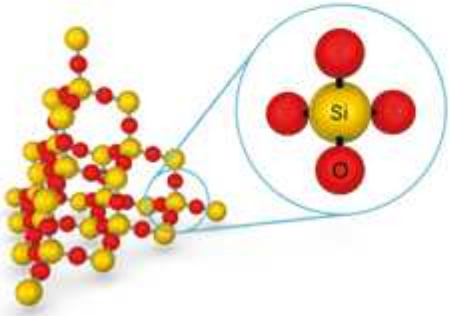


Figure 2.51: Chemical structure of silica

Science understanding

A molecule can be represented by a molecular formula.

A continuous covalent substance is represented by an empirical formula.

Molecules: Molecular formula

As molecules are discrete units comprised of a fixed number of atoms each molecular compound can be represented by a **molecular formula**. A molecular formula consists of the atomic symbols of each of the constituent elements, followed by numeric subscripts describing the number of atoms of each element present in the molecule. The central atom is written first, followed by the bonded atoms.

Naming molecules

Molecules are named systematically:

1. The central atom (usually of lowest electronegativity and furthest to the left of the periodic table) is first identified and named.
2. The bonded atom is then identified and named, ensuring the suffix is replaced with "ide".
3. Prefixes derived from Greek stems (shown in Table 2.15) are used to specify the number of atoms in a molecule.
4. If mono- is the prefix on the central atom, it is assumed and not written.

Number of atoms present	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Table 2.15: Numerical prefixes for naming molecules

Examples of molecular formulae and their systematic names are shown below in Table 2.16

Molecular formula	Systematic molecular name
CO	carbon monoxide
HCl*	hydrogen chloride
NO	nitrogen monoxide
CO_2	carbon dioxide
H_2O*	dihydrogen monoxide
KrF_2	krypton difluoride
SO_3	sulfur trioxide
CCl_4	carbon tetraoxide
N_2O_5	dinitrogen pentaoxide
SF_6	sulfur hexafluoride
IF_7	Iodine heptafluoride
As_4O_{10}	tetraarsenic decoxide

Table 2.16: Numerical prefixes for naming molecules

*except for water and acids, hydrogen is not listed first in a molecular formula.

Certain molecules were discovered and named prior to the introduction of the systematic method for deducing a molecules chemical name. Many of these molecules have maintained this name in preference to their chemical name. Examples of these "common" names are shown below.

Molecular formula	Systematic molecular name	Common name
H_2O	dihydrogen oxide	water
NH_3	nitrogen trihydride	ammonia
PH_3	phosphorous trihydride	phosphine
NO	nitrogen monoxide	nitric oxide
N_2O	nitrogen dioxide	nitrous oxide

Table 2.17: Examples of "common" molecule names

Covalent networks: Empirical formulae

Unlike molecules, the number of atoms in covalent networks are not fixed. The number of atoms present will depend on the size of the structure. The smallest amount of a network solid that can be identified is called a formula unit. A formula unit is an example of an empirical formula, denoting the simplest ratio of elements in the compound. The empirical formula of a covalent network is often contained within brackets followed by a subscript *n* as a placeholder for any number of repeating units. Common examples of covalent networks are shown in Table 2.18.

Covalent Network	Empirical formula
diamond	$(C)_n$
graphite	$(C)_n$
graphene	$(C)_n$
silica	$(SiO_2)_n$
silica carbide	$(SiC)_n$
boron carbide	$(B_4C)_n$

Table 2.18: Empirical formulae of covalent networks

Questions

14. Molecules are discrete units comprised of a fixed number of atoms. Write the molecular formula for the following molecules using their systematic names.

Systematic molecular name	Molecular formula
chlorine dioxide	
oxygen difluoride	
sulfur tetrachloride	
diphosphorus pentoxide	
sulfur dichloride	
bromine pentafluoride	
dinitrogen trioxide	
xenon trioxide	
sulfur hexafluoride	
iodine dichloride	
carbon monoxide	
dinitrogen pentasulfide	

(12 marks) **KA4**

15. Use the molecular formulae for the molecules listed in the table to write their systematic name.

Molecular formula	Systematic molecular name
ClO_2	
P_4S_5	
N_2S	
SCl_4	
B_2Si	
NF_3	
HI	
PI_3	
IF_5	
N_2O_3	
SO_3	
CS_2	

(12 marks) **KA4**

Science understanding

The physical properties of continuous covalent substances can be explained using the model for covalent bonding.

Hard solids

Covalent network solids are hard materials. Covalent bonds are extremely strong, therefore in this bonding arrangement solids require an extremely high amount of energy to break. This allows covalent network solids to withstand high degrees of mechanical stress. Diamond is the hardest known substance. A cutting instrument comprised of diamond is required to cleave diamond and form it into desirable shapes.

Science as a human endeavour:

Carbide compounds

As covalent networks, silicon carbide and boron carbide are hard ceramic materials. They have numerous societal applications. Silicon carbide can be used as an abrasive material in the manufacture of automotive brake and clutch parts. It is also the main component of bullet proof vests and body armour. Boron carbide's hardness allows it to be used as armour plating in armoured vehicles and military aircraft.

Melting and boiling point

Due to the strong nature of the covalent bonding in covalent network solids, they are solids of extremely high melting and boiling points. A significant quantity of thermal energy is required to disrupt the repeating interactions within the covalent lattice to a great enough extent to break these bonds. Diamond has a melting point of almost 4000°C.

Electrical conductivity

Most network covalent solids are poor conductors of electricity as all valence electrons are involved in covalent bonds. The material has no free electrons to carry charge to allow electrical conduction. Graphite is one of the few exceptions, as only three of the four valence electrons in the carbon atoms are involved in covalent bonding; the fourth valence electron is delocalised, and can thus move freely between the graphene layers, acting as a current carrier. This gives graphite the ability to act as an excellent conductor of electricity.

Questions

16. Naturally occurring graphite can be processed into synthetic graphite, which is useful for high value applications like lithium-ion batteries.

    (a) State and explain the physical property of graphite that makes it an appropriate component of lithium-ion batteries.

    (3 marks) **KA1**

    (b) Graphite and diamond are both allotropes of carbon, however graphite breaks and scratches easily, whereas diamond is incredibly strong.

    Explain this property with reference to the chemical structure of both graphite and diamond.

    (4 marks) **KA1**

    (c) Graphite and methane are examples of carbon-based compounds. Graphite is a solid with an extremely high melting point, methane is a gas with an incredibly low melting point.

    Explain this difference with reference to the differences in their bonding.

    (4 marks) **KA1**

17. Silica Gel Packets provide an economic and effective way of protecting against moisture damage during transit and storage. They work by absorbing moisture when there is too much and, if the air gets too dry, releasing some of the water its already absorbed. The gel itself is a mix of water and silica (silicon dioxide).

    (a) The silica contained within the gel packets has a melting point of 1,710°C.

    Explain this property with reference to silica's chemical structure.

    (2 marks) **KA1**

    (b) State and explain an additional physical property which makes silica an appropriate material to be used in silica gel packets.

    (2 marks) **KA1**

2.3: Quantities of molecules and ions

Science understanding

The number of moles of particles (molecules, ions) in a sample can be determined from the mass of the sample and the molar masses of the particles.

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Determining of the number of molecules in a sample

Chemists use the term **mole** to represent a large number of atoms or molecules. Just as a dozen implies 12 items, a mole (n) is the amount of a material containing 6.02 × 10^{23} **elementary particles**. This value is referred to as **Avogadro's number** (N_A). The number of molecules (N) in a sample can be calculated when the moles present in the sample is known using the following mathematical relationship:

$$N = n \times N_A$$

The number of molecules present in a sample will be equal to the number of moles multiplied by Avogadro's number. This sequential relationship is depicted in Table 2.19 below.

n	N_G	N
1	6.02×10^{23}	6.02×10^{23}
2	6.02×10^{23}	1.20×10^{24}
5	6.02×10^{23}	3.01×10^{24}
0.5	6.02×10^{23}	3.01×10^{23}

Table 2.19: Determining the number of particles in a compound through the multiplication of moles and Avogadro's number.

Example

Determine the number of molecules present in 0.5 mol of water.

$N = n \times N_A$

$N = 0.5 \times 6.02 \times 10^{23}$

$N = 3.01 \times 10^{23}$ molecules

Importantly, this value represents that there are 3.01×10^{23} molecules present in the sample of water. As each water molecule contains three atoms (H_2O), there are $3 \times (3.01 \times 10^{23})$ atoms or 9.03×10^{23} atoms present in the sample.

The number of moles is not always known, requiring first the calculation of moles using the mass (m) of the sample, and its molar mass (M).

1. Determine the molar mass (M) of the compound using the molar masses of the elements in the molecular or empirical formula.
2. Calculate the number of moles present using the mass of the sample.

$$n = \frac{m}{M}$$

3. Use the calculated values to determine the number of molecules present in the sample.

$$N = n \times N_A$$

Example

Determine the number of molecules present in a 22.5g sample of water.

1. Determine the molar mass (M) of the compound using the molar masses of the elements in the molecular or empirical formula.

   $M = (2 \times M_H) + M_O$

   $M = (2 \times 1.008) + (16.00)$

   $M = 18.016\ \text{g mol}^{-1}$

2. Calculate the number of moles present using the mass of the sample.

   $n = \frac{m}{M}$

   $n = \frac{22.5}{18.016}$

   $n = 1.25$ mol

3. Use the calculated values to determine the number of molecules present in the sample.

   $N = n \times N_A$

   $N = 1.25 \times 6.02 \times 10^{23}$

   $N = 7.52 \times 10^{23}$ molecules present in the sample

Example

1. Determine the number of molecules present in 102g of calcium carbonate ($CaCO_3$)

   Determine the molar mass (M) of the compound using the molar masses of the elements in the molecular or empirical formula.

   $M = M_{Ca} + M_C + (3 \times M_O)$

   $M = 40.08 + 12.01 + (3 \times 16.00)$

   $M = 100.09\ \text{g mol}^{-1}$

2. Calculate the number of moles present using the mass of the sample.

   $n = \frac{m}{M}$

   $n = \frac{102}{100.09}$

   $n = 1.02$ mol

3. Use the calculated values to determine the number of molecules present in the sample.

   $N = n \times N_A$

   $N = 1.02 \times 6.02 \times 10^{23}$

   $N = 6.14 \times 10^{23}$ molecules present in the sample

Questions

18. Our bodies synthesise protein using twenty amino acids. Determine the number of molecules in 2 g of each of the following amino acids.

(a) Methionine – $C_5H_{11}NO_2S$ M = 75.07 g mol^{-1}	(b) Glutamic acid – $C_5H_9NO_4$ M = 147.13 g mol^{-1}
(c) Tryptophan – $C_{11}H_{12}N_2O_2$ M = 204.23 g mol^{-1}	(d) Valine – $C_5H_{11}NO_2$ M = 117.15 g mol^{-1}
(e) Theanine – $C_7H_{14}N_2O_3$ M = 174.2 g mol^{-1}	(f) Glycine – $C_2H_5NO_2$ M = 75.07 g mol^{-1}

(24 marks) **KA4**

19. Cooper is commonly used to fabricate electrical wire. A length of 50 cm of copper wire weighs 7.88 g.

(a) Determine the number of moles present in the length of wire.

(2 marks) **KA4**

(b) Determine the number of atoms in the wire length.

Avogadro's number is 6.02 × 10^{23}

(2 marks) **KA4**

20. Tartaric acid or cream of tartare ($C_4H_6O_6$) is a by-product of wine production left as a residue on the inside of wine barrels. It is commonly used as a leavening agent in baking as when combined with baking soda, carbon dioxide is produced which expands the dough. Tartaric acid is sold in 125g jars.

    (a) Calculate the molar mass of cream of tartare in g mol^{-1}.

    Use the periodic table of elements to determine the molar masses of carbon, hydrogen and oxygen.

    (1 mark) **KA4**

    (b) Determine the number of moles of $C_4H_6O_6$ present in one jar of cream of tartare.

    (2 marks) **KA4**

    (c) Determine the number of $C_4H_6O_6$ molecules present in the jar of cream of tartare.

    Avogadro's number is 6.02×10^{23}

    (2 marks) **KA4**

Science understanding

The percentage composition of elements in compounds can be determined from the molar masses of the atoms.

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Percentage Composition

The number of molecules present within a sample can be analysed further through consideration of the elements which comprise it. The atomic makeup of a compound is defined by its molecular or molecular formula. The **percent composition** is described as the percent by mass of each element present in the compound. This is a useful measure to determine a substance's purity or its effectiveness for an intended purpose.

To determine percentage composition using the above mathematical relationship, the following steps can be used:

1. Identify the molecular formula of the compound in question.
2. Determine the molar mass of the compound.
3. Calculate the relative atomic mass that each atom contributes by multiplying the atom's molar mass by the number of atoms in the molecular formula. Divide this quantity by the molecules molar mass, and multiply the resultant decimal by 100 to convert it to a percentage.

This process is summarised by the following mathematical relationship:

$$\text{percentage composition } \% = \frac{M_{\text{element}} \times N_{\text{atoms of that element}}}{M_{\text{compound}}} \times 100$$

Example

Lactose ($C_{12}H_{22}O_{11}$) is the sugar component of milk produced by humans, cows, sheep, goats and other mammals.

Determine the percentage composition of each element contained within a sucrose molecule.

1. Identify the molecular formula of the compound in question.

   Lactose has the molecular formula $C_{12}H_{22}O_{11}$.

2. Determine the molar mass of the compound.

   $M = (12 \times M_C) + (22 \times M_H) + (11 \times M_O)$

   $M = (12 \times 12.01) + (12 \times 1.008) + (11 \times 16.00)$

   $M = 332.22 \text{ g mol}^{-1}$

3. Calculate the relative atomic mass that each atom contributes by multiplying the atom's molar mass by the number of atoms in the molecular formula. Divide this quantity by the molecules molar mass, and multiply the resultant decimal by 100 to convert it to a percentage.

Carbon percent composition

$$\text{percentage composition } \% = \frac{M_{\text{element}} \times N_{\text{atoms of that element}}}{M_{\text{compound}}} \times 100$$

$$\% = \frac{12.01 \times 12}{342.30} \times 100$$

$$\% = 42.10\%$$

Hydrogen percent composition

$$\text{percentage composition } \% = \frac{M_{\text{element}} \times N_{\text{atoms of that element}}}{M_{\text{compound}}} \times 100$$

$$\% = \frac{1.008 \times 22}{342.30} \times 100$$

$$\% = 6.49\%$$

Oxygen percent composition

$$\text{percentage composition } \% = \frac{M_{\text{element}} \times N_{\text{atoms of that element}}}{M_{\text{compound}}} \times 100$$

$$\% = \frac{16.00 \times 11}{342.30} \times 100$$

$$\% = 51.42\%$$

Questions

21. There are sixteen known iron oxides and oxyhydroxides, the best known of which is rust, a form of iron(III) oxide known as ferrous oxide.

    (a) Iron(II) oxide, FeO is commonly known as ferrous oxide. Determine the percentage composition of iron in the compound FeO.

(3 marks) **KA4**

(b) Determine the percentage composition of iron in the compound Fe_2O_3.

(3 marks) **KA4**

(c) State which iron oxide provides the purest source of iron.

.. ..*(1 mark)* **KA1**

22. Aspirin is a compound with the molecular formula $C_9H_8O_4$. Calculate the percent composition of each element present.

(7 marks) **KA4**

Determination of empirical formulas

The above calculations help communicate that the molecular formula of a compound represent the relative numbers, not masses of the atoms in the substance. Therefore, any experimentally derived data involving percentage composition or mass can be used to determine the corresponding numbers of atoms in a compound in their simplest ratio. When expressed in this manner it is described as a compound's empirical formula.

Deriving empirical formulas - percent composition

For instances where the percent composition data is known, it can be used along with masses of elements present in defined mass of compound to determine the empirical formula.

To determine a samples empirical formula, the following steps are used:

1. Use a defined mass (100 g) and percentage composition (represented as a decimal) for each element present in the sample to calculate the relative moles.
2. Use the relative mass and molar mass of each element present in the sample to calculate the relative moles.
3. Represent a temporary formula with each species followed by its mole value in subscript.
4. Convert the formula to accepted convention to contain whole-number subscripts by dividing each mole value by the smallest mole value.

Example

A gaseous compound produced from the cell metabolism contains 40.0% carbon, 6.71% hydrogen and 53.28% oxygen.

1. Use a defined mass (100 g) and percentage composition (represented as a decimal) for each element present in the sample to calculate the relative masses.

m_C = % composition × 100 m_H = % composition × 100 m_O = % composition × 100

$m_C = 0.40 \times 100$ $m_H = 0.0671 \times 100$ $m_O = 0.5328 \times 100$

$m_C = 40.0$ g $m_H = 6.71$ g $m_O = 53.28$g

2. Use the relative mass and molar mass of each element present in the sample to calculate the relative moles.

$$n_C = \frac{m}{M} \qquad n_H = \frac{m}{M} \qquad n_O = \frac{m}{M}$$

$$n_C = \frac{40.0}{12.01} \qquad n_H = \frac{6.71}{1.008} \qquad n_O = \frac{53.28}{16.00}$$

$n_C = 3.33$ mol $n_H = 6.66$ mol $n_O = 3.33$ mol

3. Represent a temporary formula with each species followed by its mole value in subscript.

$$C_{3.33}H_{6.66}O_{3.33}$$

4. Convert the formula to accepted convention to contain whole-number subscripts by dividing each mole value by the smallest mole value.

$$C\,\frac{3.33}{3.33}\,H\,\frac{6.66}{3.33}\,O\,\frac{3.33}{3.33}$$

Therefore, the empirical formula of the compound is CH_2O.

Questions

23. The bacterial fermentation of grain to produce ethanol forms a gas with a percent composition of 27.29% carbon and 72.71% oxygen.

Determine the empirical formula for the gas.

Empirical formula:.. (8 marks) **KA4**

24. Nicotine, an alkaloid in the nightshade family of plants that is mainly responsible for the addictive nature of cigarettes, contains 74.02% carbon, 8.710% hydrogen, and 17.27% nitrogen.

    Determine the empirical formula for the compound.

    Empirical formula: .... *(11 marks)* **KA4**

Deriving empirical formulas – mass

When the mass of each constituent is known the following steps are used to determine a samples empirical formula:

1. Use the mass and molar mass of each element present in the sample to calculate the relative moles.
2. Represent a temporary formula with each species followed by its mole value in subscript.
3. Convert the formula to accepted convention by dividing each mole value by the smallest mole value and then multiplying, if necessary, until the formula contains only whole-number subscripts.

Example

Consider a sample determined to contain 3.42 g carbon and 0.574 g hydrogen. Determine the samples empirical formula.

1. Use the mass and molar mass of each element present in the sample to calculate the relative moles.

$$n_C = \frac{m}{M} \qquad n_H = \frac{m}{M}$$

$$n_C = \frac{3.42}{12.01} \qquad n_H = \frac{0.574}{1.008}$$

$$n_C = 0.285 \text{ mol} \qquad n_H = 0.569 \text{ mol}$$

2. Represent a temporary formula with each species followed by its mole value in subscript.

   $C_{0.285}H_{0.569}$

3. Convert the formula to accepted convention to contain whole-number subscripts by dividing each mole value by the smallest mole value.

   $$C\,\frac{0.285}{0.285}\,H\,\frac{0.569}{0.285}$$

   Therefore, the empirical formula of the compound is CH_2.

Example

A gaseous compound is analysed and determined to contain 2.66 g chlorine and 4.20 g oxygen.

1. Use the mass and molar mass of each element present in the sample to calculate the relative moles.

$n_{Cl} = \frac{m}{M}$ $n_O = \frac{m}{M}$

$n_{Cl} = \frac{2.66}{35.45}$ $n_O = \frac{4.20}{16.00}$

$n_{Cl} = 0.075$ mol $n_O = 0.262$ mol

2. Represent a temporary formula with each species followed by its mole value in subscript.

   $Cl_{0.075}O_{0.262}$

3. Convert the formula to accepted convention to contain whole-number subscripts by dividing each mole value by the smallest mole value.

   $Cl\frac{0.075}{0.075}O\frac{0.262}{0.075}$

   Therefore, the empirical formula of the compound is $ClO_{3.5}$.

   In this example the empirical formula still contains a decimal. By multiplying each value by two the ratio is maintained and the final empirical formula, Cl_2O_7 is determined.

Questions

25. A sample of the black mineral hematite, an oxide of iron found in many iron ores, contains 69.94 g of iron and 30.06 g of oxygen.

    Determine the empirical formula of hematite.

Empirical formula: .... *(6 marks)* **KA4**

Summary Test 2: Combining Atoms

1. Aluminium is the most abundant metal in the Earth's crust.

   (a) Explain using the metallic bonding model how aluminium forms a solid.

....

....

....

.... *(2 marks)* **KA1**

The boiling points of three metals – lithium, potassium and aluminium – are given below.

Metal	Boiling point (°C)
lithium	1330
potassium	759
aluminium	2470

   (b) State the group, period and block these metals are found on the periodic table.

Metal	Group	Period	Block
lithium			
potassium			
aluminium			

(9 marks) **KA2**

   (c) Write the electron configuration for each of the metals.

Metal	Electron Configuration
lithium	
potassium	
aluminium	

(6 marks) **KA2**

   (d) Compare the valency of the three metals.

....

....

....

.... *(2 marks)* **KA1**

   (e) Use the metallic bonding model to explain:

      (i) Lithium's higher boiling point in comparison to potassium.

....

....

....

.... *(2 marks)* **KA1**

      (ii) Aluminium possesses a significantly higher boiling point compared to lithium and potassium.

....

....

....

.... *(2 marks)* **KA1**

(f) Aluminium is often used as a material in the production of cookware. Stainless steel is another example of a material used in this industry.

(i) State three properties that make these materials suitable for use in cookware.

(3 marks) **KA1**

(ii) Explain why stainless steel is stronger than aluminium.

(2 marks) **KA1**

(iii) Aluminium and stainless steel are both electrical conductors. Explain this property in reference to the metallic bonding model.

(2 marks) **KA1**

2. Bath salts have long been used as an easy and inexpensive way to treat mental and physical health ailments. The most popular bathing salts are Epsom salts (magnesium sulfate) and magnesium flakes (magnesium chloride), both of which contain magnesium compounds in different forms.

(a) Magnesium chloride has the molecular formula, $MgCl_2$.

(i) State the valency of the magnesium atom.

(1 mark) **KA2**

(ii) Describe the formation of the Mg^{2+} cation.

(2 marks) **KA1**

(iii) Describe the change to magnesium's atomic radius when it becomes an ion.

(2 marks) **KA1**

(b) Magnesium flakes are an example of an ionic solid. Describe how a solid results from the interactions of cations and anions.

(3 marks) **KA1**

(c) Epsom Salts can be sourced from naturally occurring springs. However these sources can be exposed to man-made pollutants and heavy metals which can contaminate the resulting salts, thus are commonly synthetically produced.

(i) Write the empirical formula for magnesium sulfate.

(1 mark) **KA2**

(ii) Using sulfate as an example, explain the term polyatomic ion.

(2 marks) **KA1**

(d) Bath salts are hard and brittle solids.

(i) Use the ionic bonding model to explain this physical property.

(2 marks) **KA1**

(ii) Describe and contrast the magnitude of magnesium chloride's melting point to that of magnesium sulfate.

(2 marks) **KA1**

(iii) Magnesium chloride has low compressive strength. Explain what is meant by this statement.

(1 mark) **KA1**

(e) As bath salts both Epsom salts and magnesium flakes are soluble. The resulting solution is termed an electrolyte.

(i) Explain what is meant by the term electrolyte.

(2 marks) **KA1**

(ii) Describe and contrast the electrical conductivity of a solid ionic compound and electrolyte solution.

(3 marks) **KA1**

3. Carbon tetrachloride was used extensively as refrigeration fluid and a propellant for aerosol cans. Gases such carbon tetrachloride have proven to have negative impact on the ozone layer, thus the production of these chemicals has been significantly phased out.

   (a) Write the electron configuration for both elements in carbon tetrachloride.

   (4 marks) **KA2**

   (b) State carbon tetrachloride's molecular formula.

   (1 mark) **KA2**

   (c) Draw the electron dot diagram for carbon tetrachloride.

   (2 marks) **KA2**

   (d) Carbon tetrachloride is a molecule comprised of polar covalent bonds.

      (i) Carbon is less electronegative than chlorine and therefore is the central atom. Explain why carbon is less electronegative.

      (2 marks) **KA1**

      (ii) Explain the formation of the single covalent bonds that form between each chlorine atom and the central carbon atom.

      (3 marks) **KA1**

      (iii) Explain why molecules of carbon tetrachloride are described as discrete units.

      (2 marks) **KA1**

      (iv) Draw a Lewis structure diagram to depict carbon tetrachloride's polar bonds, including any partial charges that are present.

      (2 marks) **KA2**

4. Silicon carbide, also known as carborundum is the main component of bullet proof vests and body armour. Its structure is very similar to diamond, with every second carbon replaced by silicon.

   (a) Identify the class of solid that silicon carbide belongs to.

   (1 mark) **KA2**

   (b) Write the formula unit of silicon carbide.

   (1 mark) **KA2**

(c) Silicon carbide is an incredibly hard material. Explain this property in relation to its bonding and structure.

(2 marks) **KA1**

(d) State and explain another physical property of silicon carbide that makes it an appropriate component of body armour.

(2 marks) **KA1**

5. Vitamin C ($C_6H_8O_6$) is an antioxidant and essential nutrient that has many functions including immune support, collagen production and reduction of oxidative stress.

(a) Vitamin C tablets contain 500 mg of Vitamin C, with each bottle containing 100 tablets.

(i) Calculate the molar mass of vitamin C in g mol^{-1}.

Use the periodic table of elements to determine the molar masses of carbon, hydrogen and oxygen.

(1 mark) **KA4**

(ii) Determine the number of moles of Vitamin C molecules present in one bottle of Vitamin C tablets.

(2 marks) **KA4**

(iii) Determine the number of molecules of Vitamin C present in one bottle of Vitamin C tablets.

Avogadro's number is 6.02×10^{23}

(2 marks) **KA4**

(b) Another compound involved in reducing oxidative stress in the body contains 40.0% carbon, 6.71% hydrogen and 53.28% oxygen. Determine the samples empirical formula.

Empirical formula: (6 marks) **KA4**

Topic 3: Molecules

3.1: Molecule polarity

Shapes of molecules

Science understanding

> The shapes of molecules can be explained and predicted using three-dimensional representations of electrons as charge clouds, and using valence shell electron pair repulsion (VSEPR) theory.
>
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The shape of a molecule is determined by the spatial arrangement of electron pairs surrounding a central atom. The arrangement of electron pairs around an atom or molecule can be explained using Valence Shell Electron Pair Repulsion (VSEPR) theory.

VSEPR theory describes that atoms within molecules are arranged into positions which minimise repulsion around the central atom. Non-bonded electron pair to non-bonded electron pair exhibits the greatest repulsion, bonded electron pair to bonded electron pair experiences the least repulsion. Electron pairs are arranged into a shape which minimises repulsion around the central atom.

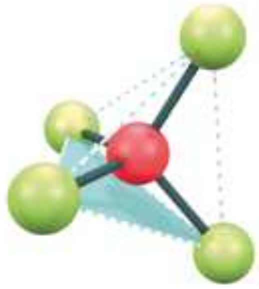


Figure 3.01: Tetrahedral molecular shape

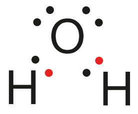
When drawing molecules, their shape is determined by following the process described below:

1. Identify the central atom (atom with the highest covalence) and draw the corresponding Lewis dot structure (discussed on page 68). Covalence is the maximum number of electron pairs that are shared by an atom.
2. Convert the electron pairs that are shared between atoms into bonds (a line) and draw any non-bonding electrons (dots).
3. Draw the 3-dimensional structure of the molecule.

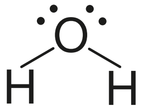
   Consider a molecule with four electron regions. The electron regions will point towards the vertices of a tetrahedron as seen in Figure 3.01.

Examples

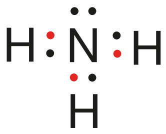
Water has two bonding electron pairs (between the O and H atoms) as well as two non-bonding pairs of electrons.



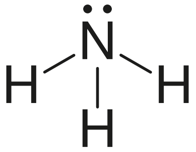
The four electron pairs are arranged into a shape which minimises repulsion around the central oxygen atom. The resulting shape is termed a **V-shaped molecule** according to VSEPR theory.



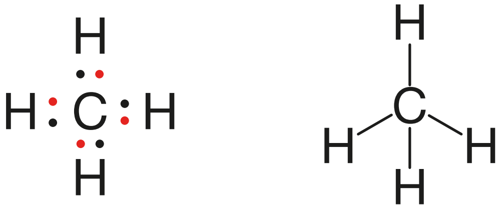
Ammonia has three bonding electron pairs (between the N and H atoms) as well as one non-bonding pair of electrons).



The four electron pairs are arranged into a shape which minimises repulsion around the central nitrogen atom and the resulting structure is termed a **trigonal pyramidal molecule**.



Methane has four bonding electron pairs which are arranged into a tetrahedron to minimise repulsion around the central carbon atom. The resulting structure is termed a **tetrahedral molecule**.

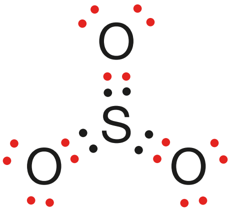


There are four bonding electron pairs around the central carbon atom within a molecule of carbon dioxide. These electron pairs are involved in two double covalent bonds between the carbon and oxygen atoms. To minimise repulsion around the central carbon atom, the bonding electron pairs are positioned 180° apart. The resulting structure is termed a **linear molecule**.

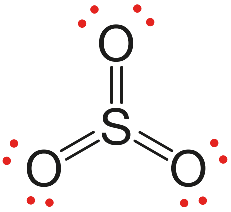
Expansion of the octet

The third electron shell of an atom has the ability to hold 18 electrons. The atoms of the chemical elements in groups V-VII of period 3 i.e. phosphorus, sulfur and chlorine, are able to expand their octet when forming molecules. These elements have the ability to share all of their valence electrons and thus acquire more than eight electrons in their valence shell.

Atoms of sulfur have six valence electrons. An atom of sulfur can combine chemically with three atoms of oxygen to form a sulfur trioxide molecule, SO_3. The central sulfur atom shares all six of its valence electrons with the three atoms of oxygen. This results in sulfur acquiring twelve electrons in the valence shell.



There are six bonding electron pairs within a molecule of sulfur trioxide. These electron pairs are involved in three double covalent bonds between the sulfur and oxygen atoms. To minimise repulsion around the central sulfur atom the bonding electron pairs are arranged into a **trigonal planar** structure.



Common molecular shapes are summarised in the following table.

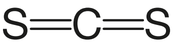
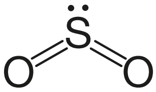
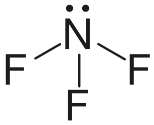
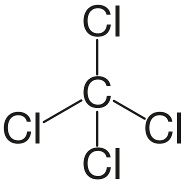
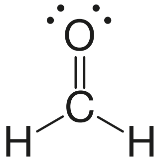
Molecular shape	Example
Linear	CS_2 
V-shaped	SO_2 
Trigonal pyramidal	NF_3 
Tetrahedral	CCl_4 
Trigonal planar	CH_2O 

Table 3.01: Summary of molecular shapes

Questions

1. Draw the dot structures for the following molecules.

(a) silane, SiH_4

(c) hydrogen sulfide, H_2S

(b) fluorine, F_2

(d) phosphine, PH_3

(4 marks) **KA4**

2. Draw the structural formula for the following molecules and name their shapes.

   (a) silane, SiH_4

   (c) hydrogen sulfide, H_2S

   (b) fluorine, F_2

   (d) phosphine, PH_3

(8 marks) **KA4**

3. An atom of phosphorous is able to expand its octet.

   (a) Explain what expansion of the octet means.

(2 marks) **KA1**

   (b) Explain why phosphorous is able to expand its octet.

(2 marks) **KA1**

Molecule polarity

Science understanding

> The polarity of a molecule results from the polar character of the bonds and their spatial arrangement.
>
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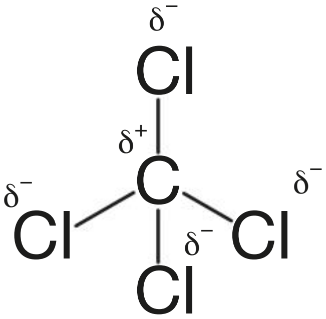
A molecule is polar if there is an uneven distribution of electric charge around the molecule. The polarity of a molecule is determined by:

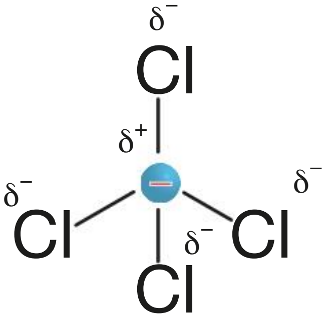
1. Identifying the shape of the molecule.
2. Determining the polarity of all the covalent bonds within the molecule.

One approach to determining the overall polarity of the molecule is to identify centres of positive and negative electric charge within the molecule. If the centres of positive and negative charge are in the same location, the electric charge distribution is even and the molecule has no electric dipole, it is then termed a **non-polarised molecule** or a **non-polar molecule**.

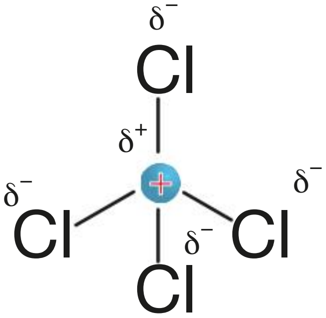
Consider a molecule of carbon tetrachloride (CCl_4). The molecule has a tetrahedral shape and the covalent bonds between carbon and chlorine atoms are polar. The uneven distribution of electrons due to the electronegativity difference between carbon and chlorine atoms results in each chlorine atom acquiring a partial negative electric charge (δ^-) and the carbon atom acquires a partial positive electric charge (δ^+).



The centre of negative charge on the carbon tetrachloride molecule is positioned at the centre of the four chlorine atoms which is found on the carbon atom.



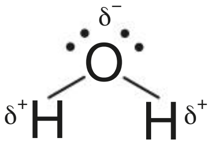
The carbon atom is the only positive charge in the molecule and, hence, the centre of positive charge on the carbon tetrachloride molecule is also positioned on the carbon atom.



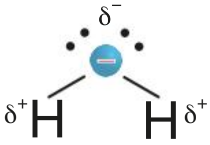
The centres of positive and negative charge are in the same location within the carbon tetrachloride molecule. This means that electric charge is evenly distributed in a molecule of carbon tetrachloride which results in the molecule having no electric dipole. A carbon tetrachloride molecule is therefore a non-polar molecule.

If the centres of positive and negative charge are in different locations, the molecule is termed a **polarised molecule** or **polar molecule**.

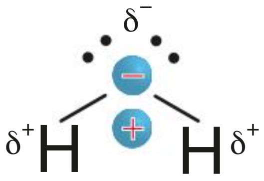
Consider a molecule of water (H_2O). The molecule has a V-shape and the covalent bonds between oxygen and hydrogen atoms are polar. The uneven sharing of electrons due to the electronegativity difference between oxygen and hydrogen atoms results in the oxygen atom acquiring a partial negative electric charge (δ^-) and each hydrogen atom acquires a partial positive electric charge (δ^+).



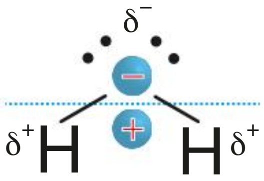
The oxygen atom is the only negative charge in the molecule and, hence, the centre of negative charge on the water molecule is positioned on the oxygen atom.



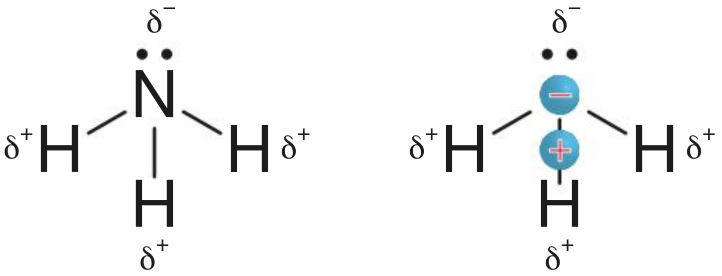
The centre of positive charge in the water molecule is positioned between the two hydrogen atoms.



The centres of positive and negative charge are in different locations within the molecule of water. This means that electric charge is unevenly distributed in a molecule of water which results in the molecule having distinct positive and negative poles. A water molecule is an electric dipole and is therefore a polar molecule.



A molecule of ammonia contains three polar covalent bonds between nitrogen and hydrogen atoms. The central nitrogen atom has a partial negative charge and each hydrogen atom has a partial positive charge. The centre of negative charge is located on the central nitrogen atom, while the centre of positive charge is located at the centre of the three hydrogen atoms.



As the centres of electric charge are in different locations within the molecule of ammonia, the electric charge is unevenly distributed. This results in the molecule having distinct positive and negative poles; ammonia is a polar molecule.

Questions

4. Draw the structural formula for the following molecules and show the polarities of the bonds using the appropriate conventions.

(a) chloromethane, CH_3Cl

(c) carbon dioxide, CO_2

(b) hydrogen sulfide, H_2S

(d) hydrogen cyanide, HCN

(8 marks) **KA1**

5. State and explain the polarity of the above molecules, using appropriate diagrams..
   (a) chloromethane, CH_3Cl

   (b) hydrogen sulfide, H_2S

   (c) carbon dioxide, CO_2

   (d) hydrogen cyanide, HCN

(12 marks) **KA4**

6. The polarity of a molecule is not only dependent on the individual bonds, but also on the geometry of the molecule. Explain using appropriate diagrams why the SO_2 molecule is polar while the CO_2 molecule is non-polar.

(6 marks) **KA4**

7. Consider the elements potassium and phosphorous.

(a) Write the formula for each chloride

potassium chloride:

phosphorous trichloride: (2 marks) **KA4**

(b) Contrast the primary bonding exhibited in the formation of these chlorides, relating your answer to:

(i) Their position on the Periodic Table

(ii) The number of valence electrons

(4 marks) **KA1**

(c) Draw a diagram showing the structural formula for the phosphorous trichloride molecule.

(2 marks) **KA4**

(d) Explain why this molecule is polar.

(4 marks) **KA1**

3.2: Interactions between molecules

Science understanding

The physical properties of molecular substances can be explained by considering the nature and strength of the forces of attraction between the molecules.

Secondary interactions between molecules are much weaker than primary metallic, ionic, and covalent bonds.

The shape, polarity, and size of molecules can be used to explain and predict the nature and strength of secondary interactions.

Physical properties (melting point/boiling point) of molecules are determined by a range of factors such as molecular size and polarity.

Primary bonds are **intramolecular** forces and secondary interactions are **intermolecular** forces. Intermolecular forces or secondary interactions are found **between** molecules or atoms and are significantly weaker than the primary bonds found **within** molecules and compounds (metallic, ionic, covalent).

Questions

8. State whether the type of forces exhibited between particles within the following compounds are intermolecular or intramolecular forces.

(a) Copper metal, Cu

(b) Diamond, C

(c) Magnesium chloride, $MgCl_2$

(d) Hydrogen gas, H_2

(4 marks) **KA1**

Dispersion forces

Science understanding

Dispersion forces exist between all molecules. Their strength depends on the size and shape of the molecules.

Dispersion forces are weak forces of attraction which exist between most atoms and molecules. Dispersion forces come about as a result of temporary dipoles.

Electrons are in constant motion around the nuclei of atoms and molecules. There is a high probability of electrons becoming unsymmetrically distributed around an atom or molecule. When this occurs, the atom or molecule temporarily becomes an electric dipole. This can occur within multiple molecules at the same time, where multiple molecules develop an electric dipole. During the short amount of time that two molecules have a temporary negative charge on one end of the molecule and a temporary positive charge on the other end, they are able to attract each other. The forces of attraction that exist between these molecules as a result of the temporary dipoles are called **dispersion forces**. This force of attraction is the weakest form of attraction between molecules and thus only a small amount of thermal energy is required to overcome the secondary interactions between the molecules.

Referring to group VI elements, halogens, as the period number increases, the molecules get larger as a result of more subatomic particles, more electron shells and therefore more electrons. Using the specific comparison of fluorine (F_2) to bromine (Br_2) we notice that at room temperature fluorine is a gas while bromine is a liquid. The diatomic halogen molecules are all non-polar given that the bonding atoms have the same electronegativity and share electrons evenly. Therefore, both diatomic molecules exhibit dispersion forces as the secondary force of attraction acting between molecules. Molecules of bromine have 70 electrons orbiting their nuclei compared with molecules of fluorine which have 18 electrons orbiting their nuclei. Polarisation of electrons is greater in a molecule of bromine, which results in a stronger electric dipole. The electric force (dispersion force) is proportional in magnitude to the charge on the molecule. As molecules of bromine form stronger dispersion forces, the molecules are drawn closer together, forming a liquid. More thermal energy is required to overcome the stronger secondary forces of attraction between bromine molecules and, hence, the boiling point of bromine is higher than that of fluorine.

Halogen	Melting point (°C)	Boiling point (°C)
Fluorine, F_2	–220	–188
Chlorine, Cl_2	–101	–35
Bromine, Br_2	–7	59
Iodine, I_2	114	184
Astatine, At_2	302	337

Table 3.02: Melting and boiling points of the halogens

Non-polar molecules can also experience temporary induced dipoles. This occurs as a result of atoms or molecules entering closer proximity to one another and the repulsive forces between negatively charged electrons from both molecules. This results in an uneven distribution of electrons around the nuclei of the atoms within a molecule and a **temporary induced dipole** within the molecule. Thus, a slightly negative charge on the side of the molecule with the greater amount of electrons and a slightly positive charge on the side with fewer electrons. Once this molecule moves next to molecule 2, the electrons from molecule 2 are attracted to the *temporary* positive charge from the original molecule, thus creating an induced dipole.

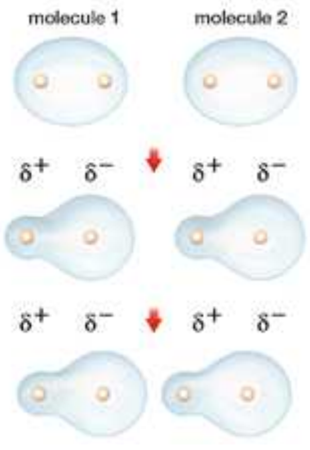
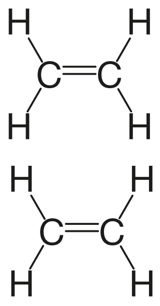
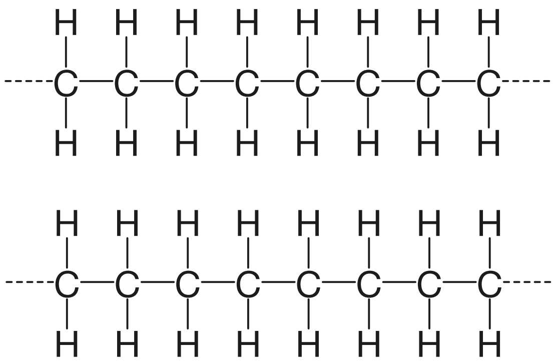


Figure 3.02: Induced dipole

Compare molecules of ethene and polyethene to one another where polyethene is a solid at room temperature while ethene is a gas.



ethene



polyethene

Both ethene and polyethene are non-polar molecules due to electric charge being evenly distributed across both molecules. Therefore, both molecules exhibit dispersion forces as the secondary forces of attraction acting between them. Polyethene has more ‘sites’ for the polarisation of electrons to occur and thus more dipole moments. Polarisation of electrons is greater in a molecule of polyethene and can occur more frequently, which results in a stronger electric dipole. As molecules of polyethene form stronger dispersion forces, the molecules are drawn closer together forming a solid. More thermal energy is required to overcome the stronger secondary forces of attraction between polyethene molecules and, hence, the boiling point of polyethene is much higher than that of ethene.

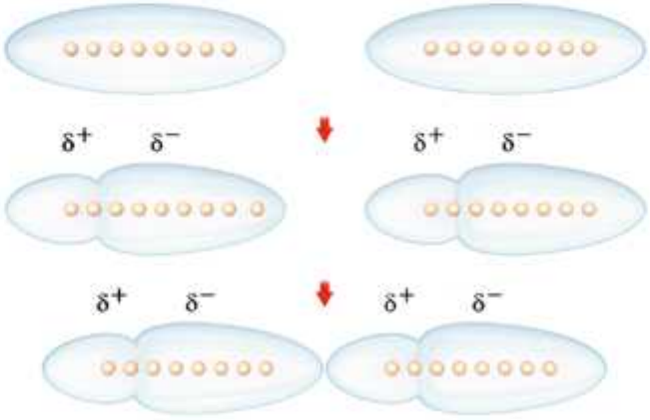


Figure 3.03: Dispersion forces acting between molecules of polyethene

The boiling point and melting point of an element can therefore be predicted by the molecular size of the element and this is a direct representation of the strength of the dispersion forces acting between molecules.

Question

9. Using your understanding of dispersion forces, predict and explain whether hydrogen gas or fluorine gas would have a higher boiling point

(3 marks) **KA2**

Dipole–dipole interactions

Science understanding

Dipole–dipole interactions exist between polar molecules and their strength depends on the polarity and size of the molecules.

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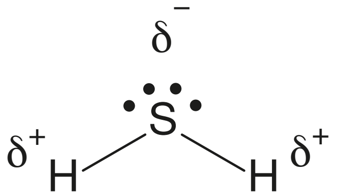
Polar molecules are electric dipoles and have distinct regions of positive and negative electric charge. Electrical forces of attraction operate between the positive charge on one polar molecule and the negative charge on a second polar molecule. The electric force of attraction between two dipoles (polar molecules) is referred to as a **dipole–dipole interaction**.

A molecule of hydrogen sulfide is shown below; delta signs (δ) are used to identify the slightly charged ends of each bond. Using the processes discussed previously in this chapter, we can determine that this is in fact a polar molecule due to the uneven distribution of electric charge within the molecule.



Molecules of hydrogen sulfide arrange themselves so that the positive charge of one molecule is aligned with the negative charge of another molecule. When these molecules are arranged this way they have formed secondary interactions called dipole–dipole interactions.

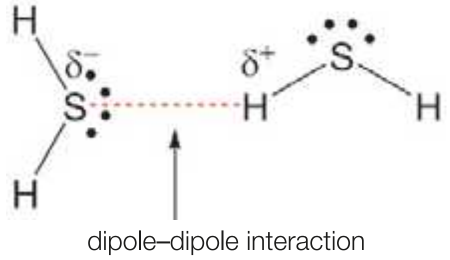


Figure 3.04: Secondary interactions between molecules of hydrogen sulfide

These secondary interactions are strong in comparison to dispersion forces and, as a result, more thermal energy is required to break the dipole–dipole interactions and separate molecules of hydrogen sulfide from each other i.e. they have a higher melting point and boiling point than molecules exhibiting dispersion forces only with similar molar mass.

As the molecular size of molecules exhibiting dipole–dipole interactions increases, the strength of the secondary interactions exhibited between these molecules also increase. If the boiling points of hydrogen bromide (–66°C) and hydrogen iodide (–34°C) are compared it can be seen that, as the period number increases in this group (VII), the boiling point increases. Both hydrogen bromide (HBr) and hydrogen iodide (HI) are polar molecules due to electric charge being unevenly distributed across both molecules. Therefore, both molecules exhibit dipole–dipole interactions as the secondary forces of attraction acting between them. Molecules of hydrogen bromide have 36 electrons orbiting their nuclei compared with molecules of hydrogen iodide which have 54 electrons orbiting their nuclei. Polarisation of electrons is greater in a molecule of hydrogen bromide which results in a stronger electric dipole. The electric force (dipole–dipole interaction) is proportional in magnitude to the charge on the molecule. As molecules of hydrogen bromide form stronger dipole–dipole interactions, more thermal energy is required to overcome the stronger secondary forces of attraction between hydrogen bromide molecules and, hence, the boiling point of hydrogen bromide is higher than that of hydrogen fluoride.

Questions

10. Hydrogen chloride (molar mass 36.46 g mol^{-1}) has a boiling point of –85.05°C. Fluorine (molar mass 38.00 g mol^{-1}) has a boiling point of –188.1°C.

    (a) State the secondary interactions that exist between molecules of hydrogen chloride.

    .. ..*(1 mark)* **KA1**

    (b) State the secondary interactions that exist between molecules of fluorine.

    .. ..*(1 mark)* **KA1**

    (c) Explain why, when both have similar molecular masses, their boiling points are considerably different using an understanding of secondary interactions.

    Use diagrams where appropriate.

    .. ..

    .. ..

    .. ..

    .. ..

    .. ..

    .. ..*(4 marks)* **KA2**

Hydrogen bonding

Science understanding

Hydrogen bonding is a particularly strong form of dipole-dipole interaction that exists between molecules.

A stronger form of dipole–dipole interaction is called **hydrogen bonding**. This occurs only between molecules that have polar covalent bonds formed between atoms of hydrogen and either nitrogen, oxygen or fluorine. The atoms of these three elements have very high electronegativity. When atoms of nitrogen (N), oxygen (O) or fluorine (F) covalently bond with atoms of hydrogen, the difference in electronegativity between the bonding atoms is large. This causes a large polarisation of electrons across the covalent bond between the atoms resulting in a stronger electric dipole. The magnitude of the positive and negative charge at each pole of the bond is large, which results in a proportional increase in the magnitude of the secondary interactions between molecules. The hydrogen bond occurs between a non-bonding electron pair of an N, O or F atom from one molecule to the slightly positively charged hydrogen atom (which is covalently bonded to an N, O or F atom). A hydrogen bond is drawn using the following convention.

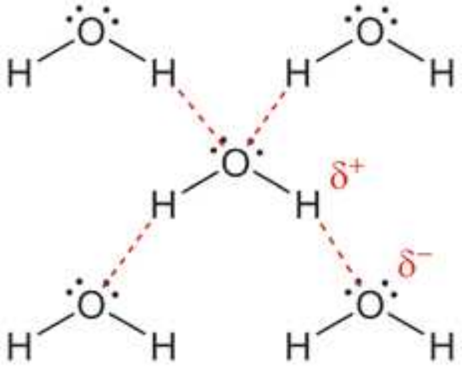


Figure 3.05: Hydrogen bonding between molecules of water

Previously we have discussed that molecular size is a contributing factor to the strength of secondary interactions. As hydrogen bonds are a lot stronger than all other secondary interactions, the boiling point of molecules exhibiting hydrogen bonding tend to be higher than expected. As an example refer to the following graph of non-metal hydrides and their corresponding boiling points.

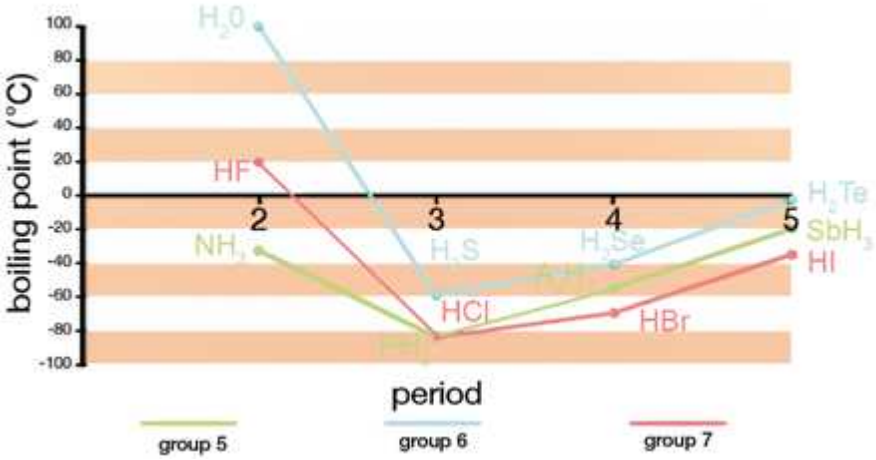


Figure 3.06: Boiling points of molecules

The boiling point of the group VI hydrides – water (H_2O) and hydrogen sulfide (H_2S) – are significantly different in magnitude. Referring to what is known about the influence of molecular size it would be expected that hydrogen sulfide would have a higher boiling point than water. However, molecular size is not the only contributing factor to the physical properties of molecules. The secondary interactions exhibited between molecules of hydrogen sulfide are dipole–dipole interactions. The secondary interactions exhibited between molecules of water are hydrogen bonds. Hydrogen bonding is a stronger form of secondary interaction and is therefore stronger than dipole–dipole interactions. Thus, more thermal energy is required to overcome the secondary forces of attraction between molecules of water.

Questions

11. Identify the type of secondary interactions exhibited between molecules of:
    (a) hydrogen, H_2
    (b) sulfur dioxide, SO_2
    (c) carbon dioxide, CO_2
    (d) hydrogen fluoride, HF

(4 marks) **KA1**

12. For the following questions indicate on the structures any polar bonds using appropriate conventions.
    (a) Draw the structural formula for water.

(2 marks) **KA4**

    (b) Draw the structural formula for ammonia.

(2 marks) **KA4**

    (c) Explain how molecules of water would interact with molecules of ammonia.
    Use diagrams to assist your answer.

(4 marks) **KA1**

(d) Explain, with the use of diagrams, why water at 0°C is less dense than water at 4°C.

(4 marks) **KA1**

3.3: Hydrocarbons

Hydrocarbons as fuels and as feedstock (raw materials)

Science understanding

Carbon forms hydrocarbon compounds, including alkanes and alkenes.

Hydrocarbons are used as fuels and as feedstock for the chemical industry.

Carbon chemically combines with hydrogen to form a category of molecules called **hydrocarbons**. Carbon is classified as the backbone to all living things and hence molecules containing a carbon backbone are called **organic molecules** and belong to a branch of chemistry called **organic chemistry**. Common hydrocarbons include methane (natural gas), propane (liquefied petroleum gas or LPG) and octane (found in petrol). All of these hydrocarbons are used as fuels due to the similar chemical properties they exhibit and, in particular, when they combust, the large amount of energy that they release. **Combustion** is a type of chemical reaction in which a fuel reacts with an oxidiser at high temperatures. Hydrocarbons are called fuels as they store chemical energy within the covalent bonds between atoms. An **oxidiser** is any chemical which gains electrons in a chemical reaction. In a combustion reaction, the oxidiser removes electrons from the covalent bonds within the hydrocarbon molecules. The removal of electrons from chemical bonds is accompanied by the release of the stored chemical energy. This energy is transformed into light and heat, which radiate away from the site of the reaction. The combustion of a hydrocarbon fuel can be **complete** or **incomplete**.

Complete combustion occur when hydrocarbons react with sufficient oxygen molecules to form carbon dioxide and water. In this case, there are enough moles of oxygen present to ensure all carbon atoms (from the hydrocarbon) are chemically combined with two atoms of oxygen to produce carbon dioxide.

$$2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(l)}$$

Incomplete combustion occurs when there are not sufficient oxygen molecules for the hydrocarbon to react with. When combustion is incomplete, there are not enough moles of oxygen present to combine with carbon to produce carbon dioxide. In this case, atoms of carbon may chemically combine with one atom of oxygen to form carbon monoxide (CO) or no atoms of oxygen to form elemental carbon (C).

$$2C_8H_{18(l)} + 17O_{2(g)} \rightarrow 16CO_{(g)} + 18H_2O_{(l)}$$

or

$$2C_8H_{18(l)} + 9O_{2(g)} \rightarrow 16C_{(s)} + 18H_2O_{(l)}$$

When hydrocarbons are burnt, they provide sufficient energy for other chemical reactions to occur. This a particular property used in chemical industry. When hydrocarbons are used in this way they are referred to as **fuels**. Feedstock is defined as a raw material used for processing or use in the manufacturing industry. As feedstock, **hydrocarbons** are used to produce pharmaceuticals, pesticides and herbicides, solvents for the paint industry and monomers for polymers. All of these products are extremely common in everyday life and are heavily relied upon by important industries such as hospitals, electronics, housing and food. As these raw materials are so heavily relied upon in today's society there needs to be alternative sources for the feedstock in chemical industry. This is because all fossil fuels are a finite resource and their supply is not replenished. Scientists are now turning to other fuel sources called **biomass fuels**, which are renewable, as an alternative to the traditional fossil fuels.

Questions

13. Write balanced equations for the complete combustion of the following hydrocarbons.

    (a) ethane, C_2H_6

    ....

    (b) propane, C_3H_8

    ....

    (c) hexane, C_6H_{14}

    ....

(6 marks) **KA4**

14. Carbon monoxide is a gas produced in combustion engines of vehicles.
    (a) State the type of combustion that produces carbon monoxide.

(1 mark) **KA1**

    (b) Write a balanced equation showing the production of carbon monoxide when octane is reacted with oxygen in combustion engines.

(2 marks) **KA4**

Hydrogen fluoride is an important feedstock used in the production of compounds in the pharmaceutical industry.

    (c) Draw a diagram showing the structural formula for the hydrogen fluoride molecule and identify its shape.

(3 marks) **KA4**

    (d) Explain why the hydrogen fluoride molecule has this shape.

(2 marks) **KA1**

Carbon monoxide has a molar mass of 28.0 g mol^{-1} and hydrogen fluoride has a similar mass of 20.0 g mol^{-1}. Even though carbon monoxide is a heavier molecule with a boiling point of –191.5ºC, it has a much lower boiling point than the smaller molecule hydrogen fluoride (19.5ºC).

    (e) Explain why, when both carbon monoxide and hydrogen fluoride have similar molecular sizes, their boiling points are so different using your understanding of secondary interactions.

(4 marks) **KA2**

Physical properties of hydrocarbons

Science understanding

The physical properties of hydrocarbons depend on the size of the molecules.

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Physical properties of hydrocarbons can be determined by the size of the molecule or the length of the hydrocarbon chain. As the length of the hydrocarbon chain increases, their boiling point and melting point also increase. The magnitude (strength) of the intermolecular forces (dispersion forces) between molecules increases with molecular size. As the hydrocarbon chain increases in length so does the non-polar portion of the molecule (due to the C-H bond being non-polar) and also the molecular size. As the number of carbon atoms within the molecule increases, so too does the number of sites (atoms) over which dispersion forces act between molecules.

Name of alkane	Melting point (°C)	Boiling point (°C)
methane	-182	-162
ethane	-183	-89
propane	-187	-42
butane	-138	0.5
pentane	-130	36
hexane	-94	69
heptane	-90	98
octane	-57	125
nonane	-54	150
decane	-30	174

Table 3.03: Melting and boiling points of hydrocarbons

Homologous series

Within organic chemistry, compounds are organised into groups depending on their properties. A **functional group** is an atom or group of atoms responsible for the typical chemical reactions of a molecule. A **homologous series** is a group of molecules with the same functional group but a different number of $–CH_2$ groups. Functional groups determine the pattern of reactivity of a homologous series, whereas the carbon chain length determines physical properties such as melting/boiling points. All members within a homologous series only vary by the length of their carbon chains and, as discussed previously, will therefore have differing physical properties, as these are dependent on the size of the molecule. The functional group present in an organic molecule determines the chemical properties and thus, as all members of a homologous series have the same functional group, they will all have very similar chemical properties.

Organic compounds from a homologous series therefore have the following general characteristics:

1. All members of a homologous series conform to a general formula.
2. Each member of a homologous series contains one more carbon atom (and therefore CH_2 group) than the previous member.
3. The physical properties of members within a series gradually change as the molecules get larger.
    (a) Melting point increases
    (b) Boiling point increases
4. The chemical properties of the members are all similar.

As all members of a homologous series adhere to the same set of rules, if one or two compounds of a series are known then the properties of other compounds within that series can be predicted.

Alkanes

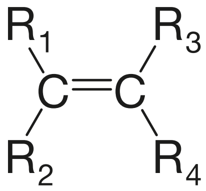
The alkanes are a homologous series of hydrocarbons which adhere to the general chemical formula C_nH_{2n+2} (where *n* is the number of carbon atoms in the molecule), and all have names ending in -ane. The first ten molecules in the homologous series of alkanes are named as follows.

Number of carbon atoms *(n)*	Molecular formula (C_nH_{2n+2})	Name
1	CH_4	methane
2	C_2H_6	ethane
3	C_3H_8	propane
4	C_4H_{10}	butane
5	C_5H_{12}	pentane
6	C_6H_{14}	hexane
7	C_7H_{16}	heptane
8	C_8H_{18}	octane
9	C_9H_{20}	nonane
10	$C_{10}H_{22}$	decane

Table 3.04: Molecular formula and names of first 10 alkanes

Alkenes

The homologous series of alkenes relates to any hydrocarbon with the functional group of a **carbon–carbon double covalent bond**. These have the general formula C_nH_{2n} and names ending in -ene.

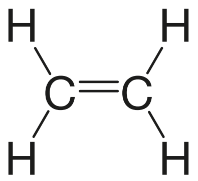


The **R** in the above diagram identifies any hydrocarbon structure or hydrogen atom.

The simplest of all the alkenes is ethene, which has the following molecular formula and structural formula.

Molecular formula: C_2H_4

Structural formula:



Alkynes

The homologous series of alkynes relates to any hydrocarbon with the functional group of a **carbon–carbon triple covalent bond**. These have the general formula C_nH_{2n-2} and names ending in -yne.

$$R_1{-}C{\equiv}C{-}R_2$$

The simplest of all the alkynes is ethyne, which has the following molecular formula and structural formula.

Molecular formula: C_2H_2

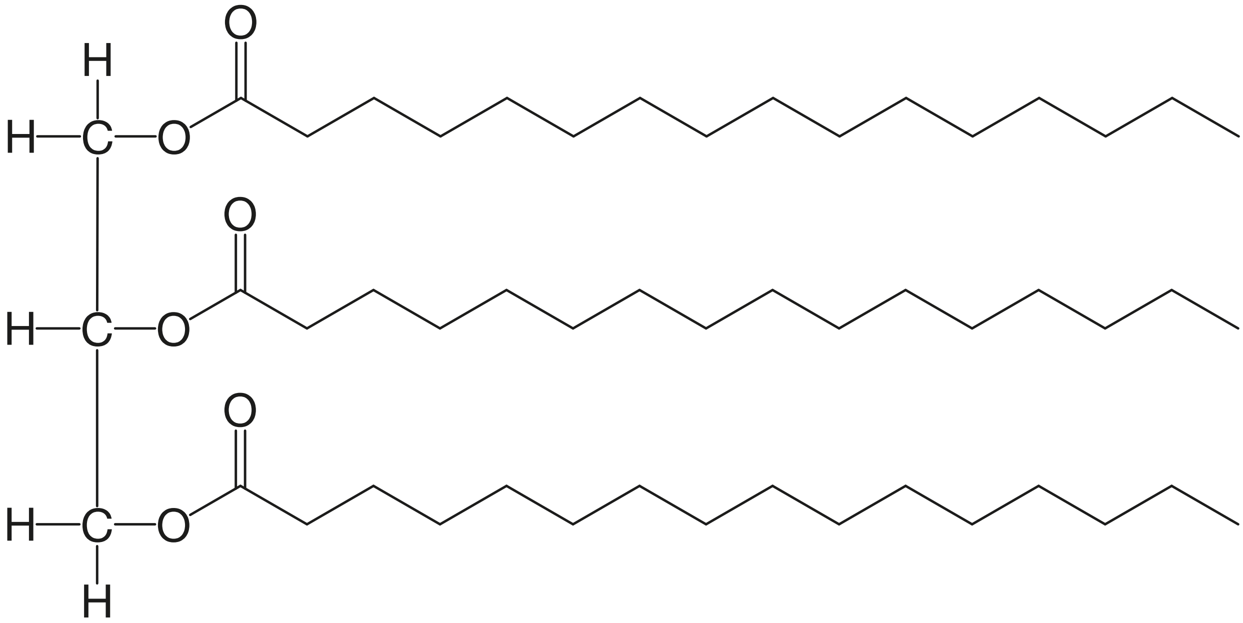
Structural formula:

$$H{-}C{\equiv}C{-}H$$

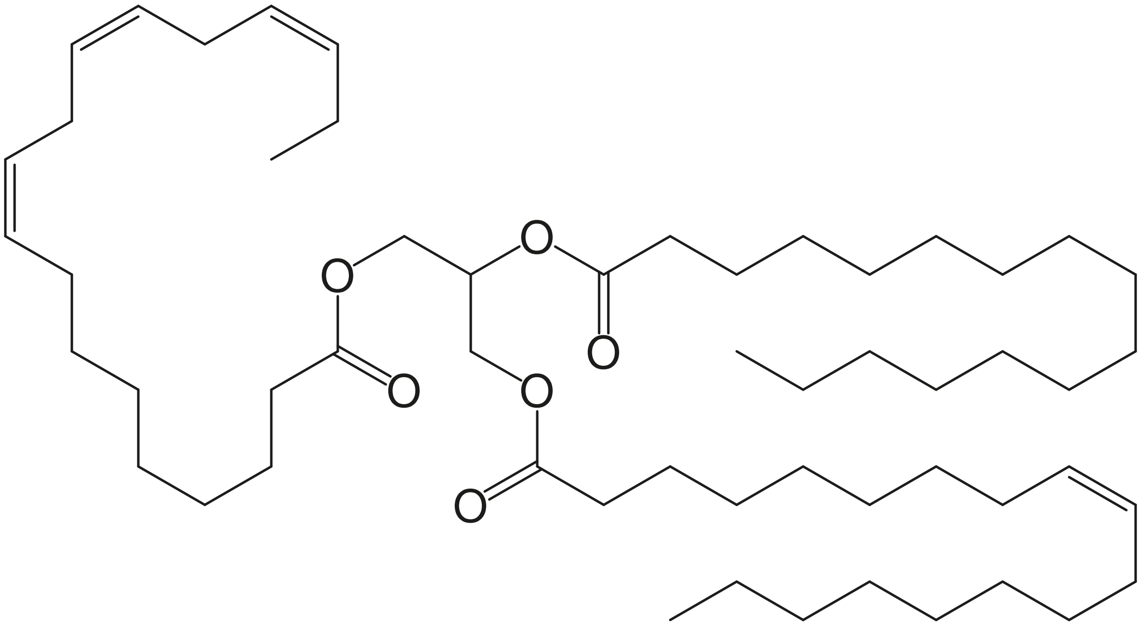
Saturated and unsaturated hydrocarbons

If an organic compound contains the maximum number of hydrogen atoms around each carbon atom that compound is said to be **saturated**. If an organic compound contains less than the maximum number of hydrogen atoms around each carbon atom the compound is **unsaturated**.

Butter is a table spread that is made from a variety of trigylcerides that are usually of animal origin and are saturated. One such molecule is glyceryl stearate (mutton/beef fat); it is a solid at room temperature.



Margarine is a spread that is made from vegetable oils. One such oil is olive oil which is made from glyceryl oleate. This oil is unsaturated and is a liquid at room temperature.



From the structures above, we can see that both are predominantly non-polar molecules due to electric charge being evenly distributed across both molecules. Therefore, both molecules exhibit dispersion forces as the secondary forces of attraction acting between them. The glyceryl stearate molecule is a straight molecule that packs and stacks with other like molecules. Glyceryl oleate, however, does not have straight chains. Therefore, when like molecules enter into closer proximity to one another, they would have fewer 'sites' (than glyceryl stearate) for the polarisation of electrons to occur and are further apart creating less dipole moments. Polarisation of electrons is greater in a molecule of glyceryl stearate and can occur more frequently, which results in a stronger electric dipole. As molecules of glyceryl stearate form stronger dispersion forces, the molecules are drawn closer together to form a solid. More thermal energy is required to overcome the stronger secondary forces of attraction between glyceryl stearate molecules and, hence, the boiling point of glyceryl stearate (animal fat) is higher than that of glyceryl oleate (vegetable oil).

Consumers may prefer to use vegetable oil as a spread as butter contains cholesterol which has been shown to increase the risk of heart disease. However, up until the 1890s this was not an option, as vegetable oils are liquid. Unsaturated vegetable oils can be made solid by converting some of the double bonds between carbon atoms into single bonds.

Science as a human endeavour:

In the 1890s, American chemist James F Boyce Senior developed a method for processing cottonseed oil into a material of greater viscosity, more suitable for use in soap and food products. He discovered that the addition of a small amount of nickel to the mixture facilitated the addition of hydrogen molecules to the product, thereby changing its properties favourably. This work was also picked up by renowned French chemist Paul Sabatier, who refined the process into the catalytic hydrogenation that earned him a Nobel Prize and today bears his name. Boyce and Sabatier were after a process that would convert less saturated (and less expensive) plant oils into molecules more closely resembling animal fats. This is because animal fats have a higher degree of saturation, making them solids near room temperature and increasing their shelf life. The double bonds of hydrocarbons can be saturated using molecular hydrogen at or near room temperature. But, importantly, the reaction is painfully slow. Boyce and Sabatier realised that the solution to this was the use of a catalyst to weaken the bond in the hydrogen molecule, thereby promoting the reaction without the need for excessive temperatures.

In the **hydrogenation** process, the oil is first heated in the presence of a nickel catalyst. Hydrogen gas is then bubbled through the hot oil under pressure. This results in the hydrogen gas ‘adding’ across the double bonds (see Figure 3.3.4 for more on addition reactions).

Polyunsaturated table spreads begin with an oil where there are some double bounds within the triglyceride structure. These are then partially hydrogenated to form a triglyceride where some but not all double bonds have been converted into single bonds. This forms a product with a similar consistency to butter (animal fat).

Chemical properties of hydrocarbons

Science understanding

The chemical reactions of hydrocarbons are determined by the functional groups present.

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Hydrocarbons undergo similar chemical reactions depending on the functional groups present within the molecule. An example of this is that alkene and alkyne molecules undergo addition reactions due to the presence of the double and triple covalent carbon-carbon bonds within the molecule. In an **addition** reaction, atoms will add to adjacent atoms either side of a double/triple covalent bond.

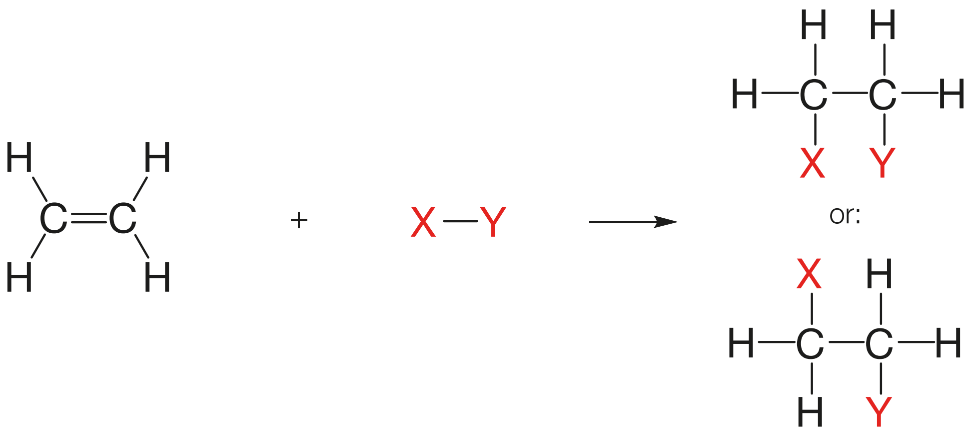

Figure 3.07: Addition reaction

Upon reacting an alkene with bromine water, the alkene will undergo an addition reaction. The double bond of an alkene contains four electrons. Atoms and molecules which have a high electronegativity will gain electrons from this double bond during a chemical reaction. The halogens (F_2, Cl_2, Br_2 and I_2) all have high electronegativities and will attack a carbon–carbon double bond. When this occurs, the halogen atoms remove electrons from the carbon–carbon double bond forming a saturated alkane. Since the alkane contains one or more halogen atoms, the molecule is termed a **halogenalkane** or **haloalkane**.

Alkenes

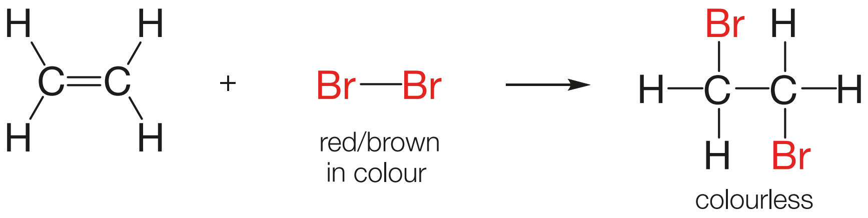

Figure 3.08: Reaction of bromine water with ethene

Alkynes

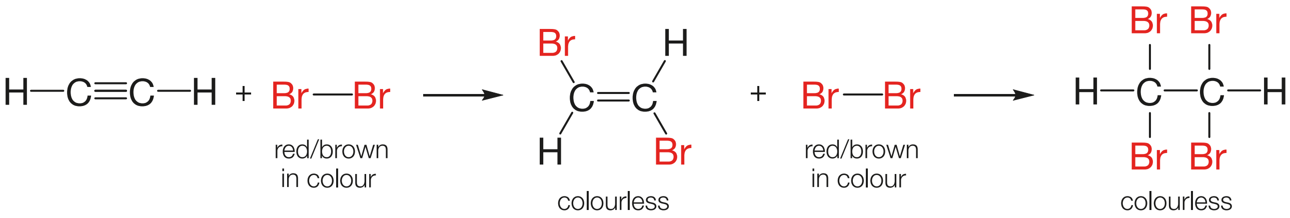

Figure 3.09: Reaction of bromine water with ethyne

This reaction is used in laboratory situations as a test to determine whether a hydrocarbon is saturated or unsaturated. Molecular bromine ($Br_{2(l)}$) is red/brown in colour. Bromine is often diluted with water forming bromine solution ($Br_{2(aq)}$), which is red/brown in colour. The test is able to be carried out because an alkane (saturated hydrocarbon) will not react with the bromine and therefore the solution will remain red/brown in colour.

Representation of hydrocarbon molecules

Science understanding

Hydrocarbons can be represented by empirical formulae, molecular formulae, and structural formulae, including extended, condensed, and skeletal representations.

Hydrocarbons can exist as different structural isomers.

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Hydrocarbon molecules can be represented in a variety of ways dependent on the way in which they are going to be used.

Molecular formulae are the most common and are used when writing chemical equations. The molecular formula of a hydrocarbon is important as it identifies the elements and the proportion of each atom in the hydrocarbon molecule.

Structural formulae can also be used for chemical equations and are graphical representations of hydrocarbon molecules. Structural formulae are particularly useful for determining the outcome of organic reactions, as most are dependent on the functional groups within the structure of the hydrocarbon. This was seen in the previous section; bromine will only react with double or triple covalent bonds.

A molecule's **empirical formula** is the smallest whole number mole ratio of the elements in a compound and it can also be the molecule's molecular formula. The empirical formula does not always define the actual number of atoms of each element in a molecule.

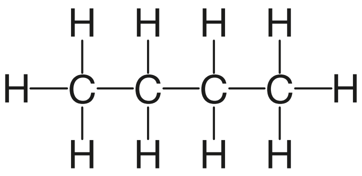
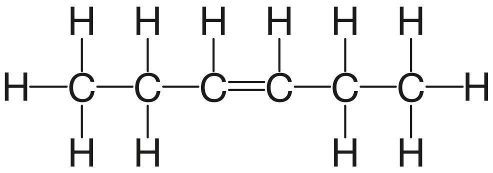
Compound	Molecular formula	Structural formula	Empirical formula
butane	C_4H_{10}		C_2H_5
hex-3-ene	C_6H_{12}		CH_2

Table 3.05: Contrast of organic formulae

The structural formula of a compound shows how the atoms are arranged in a molecule and, in particular, identifies the functional groups that are present. Other forms of structural formulae can be used to describe the molecular structure of an organic compound.

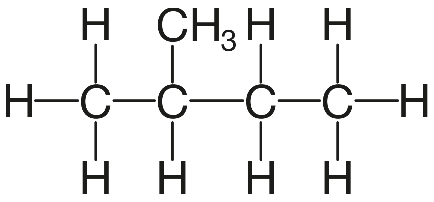
The **condensed formula** of a compound shows all of the atoms within the molecule, but not the bonds between the atoms.

$$CH_3CH_2CH{=}CHCH_2CH_3$$

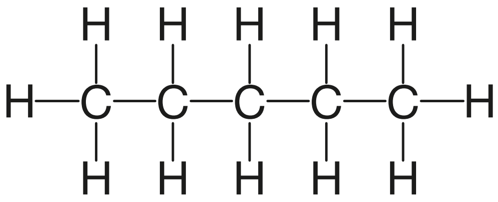
The **skeletal formula** of a compound shows the bonds between carbon atoms, but not the atoms themselves. Hydrogen atoms are also omitted, but other atoms are shown.

Structural isomers of an organic compound are that which have the same molecular formula (i.e. same number and type of atoms) but different structural formulae.

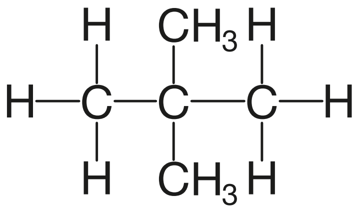
An example of this is 2-methylbutane which has the molecular formula C_5H_{12}.



The structural isomers of 2-methylbutane are given below.



Pentane is a structural isomer of 2-methylbutane, as it has the same molecular formula but a different structural formula.



2,2-dimethylpropane is a structural isomer of 2-methylbutane.

Systematic nomenclature

Science understanding

Hydrocarbons are named systematically to provide unambiguous identification.

The structural formula of a hydrocarbon can be deduced from its systematic name.

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The naming of hydrocarbons is also called systematic nomenclature and is used according to the **International Union of Pure and Applied Chemistry (IUPAC)** guidelines. IUPAC is an organisation tasked with providing internationally similar rules for naming and drawing conventions in chemistry.

Hydrocarbons are named first by determining the parent chain and naming according to the number of carbon atoms in that unbroken carbon chain. The prefix given for the number of carbon atoms can be seen in the table below.

Number of carbon atoms	Parent name prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-
9	non-
10	dec-

Table 3.06: Parent name prefixes

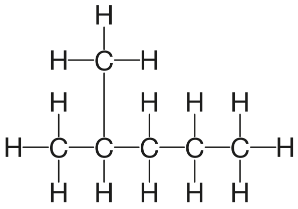
Alkanes with four or more carbon atoms display structural isomerism because the carbon chain may be either straight or branched. The naming of alkanes depends on whether they are straight or branched. **Branched hydrocarbons** contain one or more alkyl groups which "branch" from the parent hydrocarbon chain. When naming alkyl groups, they are named depending on the length of the carbon chain and receive the –yl suffix.

For example, an alkyl group of one carbon is called a methyl group, two carbons is an ethyl group, three carbons is a propyl group, etc.

Hydrocarbons are named in a systematic manner so that there is no confusion as to what the structural formula for a particular hydrocarbon molecule is. The conventions for naming structural formulae are as follows:

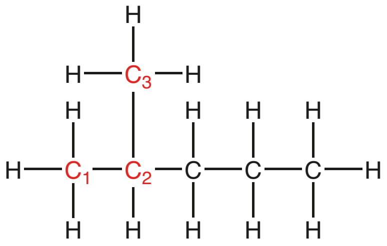
Naming a branched hydrocarbon

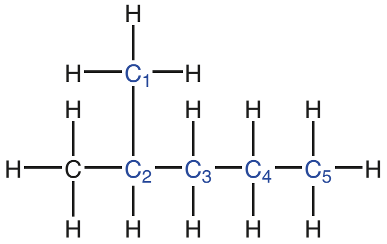
Consider the following molecule:

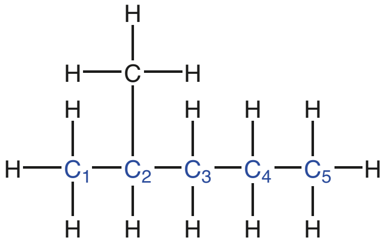


1. Identify the longest unbroken carbon chain and give this the corresponding parent name prefix.

   Note: Incorrect procedures for naming conventions are shown in red, correct procedures are shown as blue.

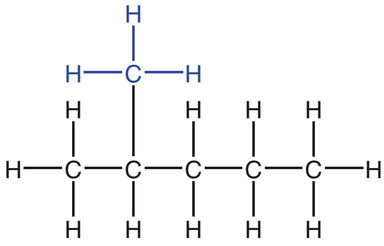






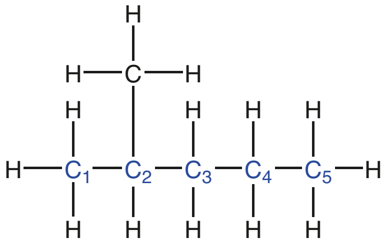
The longest unbroken carbon chain is 5 carbons in length and the molecule is saturated, so it is an alkane.

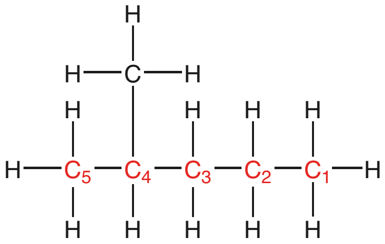
2. Locate and name the branch.



This molecule is a branched hydrocarbon where the branch is a methyl group.

3. Number each carbon atom in the parent hydrocarbon chain in a manner which results in the branch being located at the lowest-numbered position on the hydrocarbon molecule.

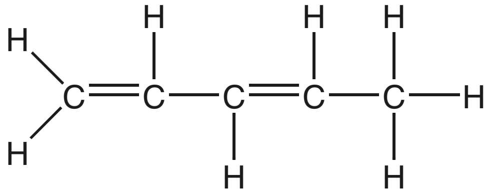




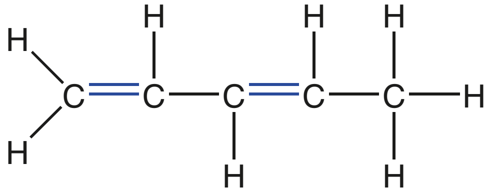
The lowest numbered position for the methyl group is on the second carbon. Therefore, this molecule is named **2-methylpentane.**

Naming a hydrocarbon which is a diene.

Consider the following molecule:

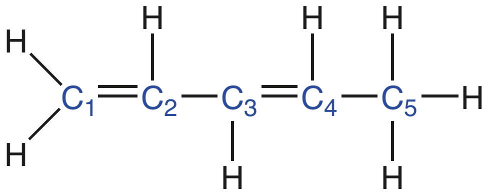


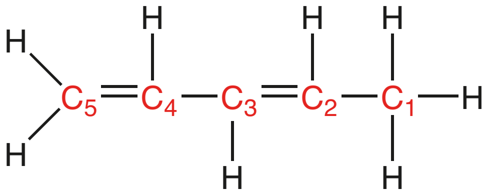
1. Locate and name the functional group(s).



This hydrocarbon is unsaturated and so is an alkene as indicated by the two C-C double bonds.

2. Number each carbon atom in the parent hydrocarbon chain in a manner which results in the functional groups being located at the lowest-numbered position on the hydrocarbon molecule.

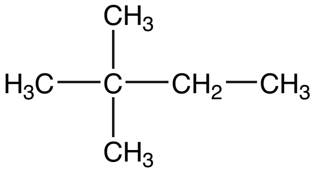




The lowest numbered position for the alkene groups is such that the double bonds are found on the first and third carbons. Therefore, this molecule is named **penta-1,3-diene**.

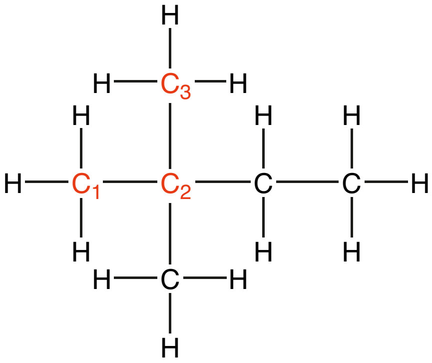
Naming a hydrocarbon which has two or more branched groups

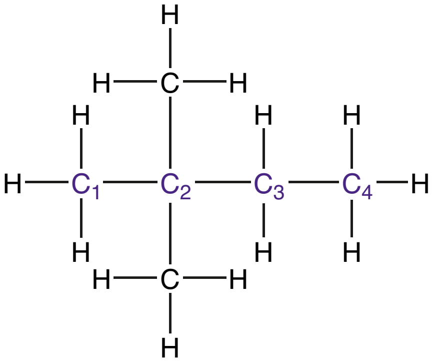
Consider the following molecule:

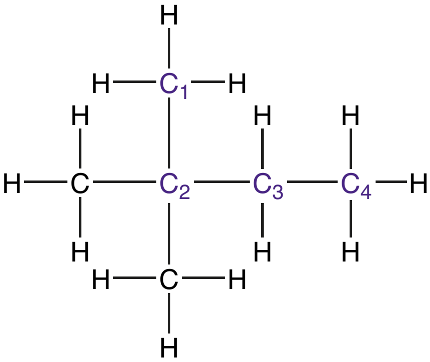


$$H_3C-\underset{CH_3}{\overset{CH_3}{C}}-CH_2-CH_3$$

1. Identify the longest unbroken carbon chain involving the key functional group and give this the corresponding parent-name prefix.

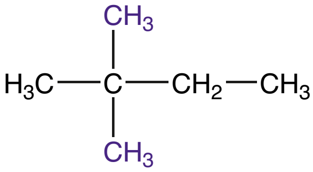






The longest unbroken carbon chain is 4 carbons in length and the molecule is saturated, so it is an alkane.

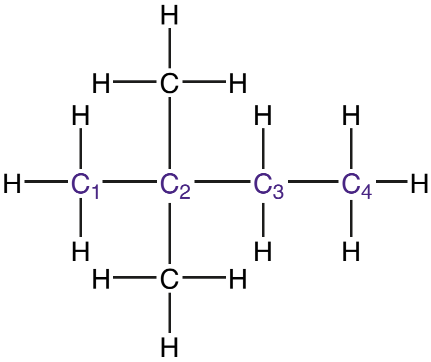
2. Locate and name the substituent group(s). If there are multiple locations for the same functional group give them the prefix di-, tri-, tetra-.

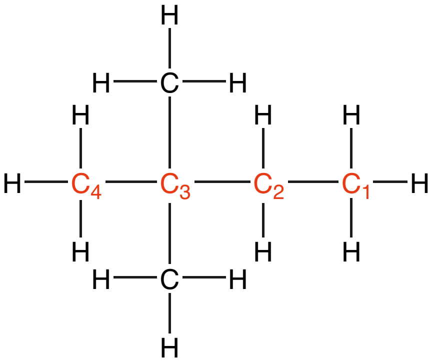


$$H_3C-\underset{CH_3}{\overset{CH_3}{C}}-CH_2-CH_3$$

This molecule is a branched hydrocarbon where there are two methyl groups.

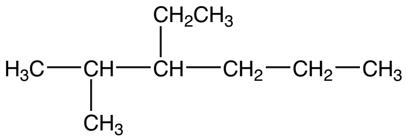
3. Number each carbon atom in the parent hydrocarbon chain in a manner which results in the substituent groups being located at the lowest-numbered position on the hydrocarbon molecule.





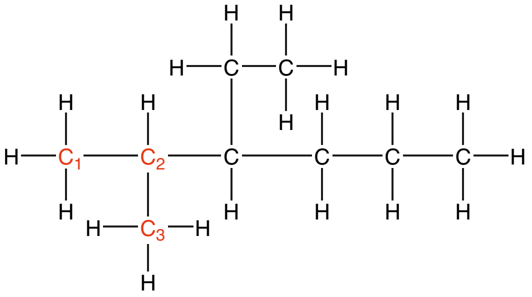
The lowest numbered position for the methyl groups is on the second carbon. Therefore, this molecule is named **2,2-dimethylbutane.**

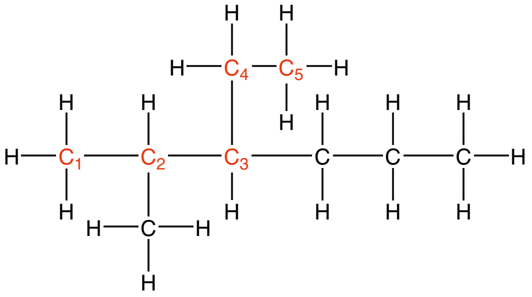
Consider the following molecule:

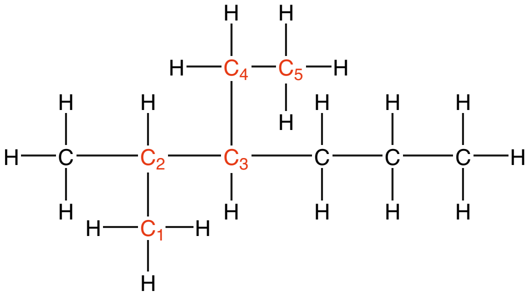


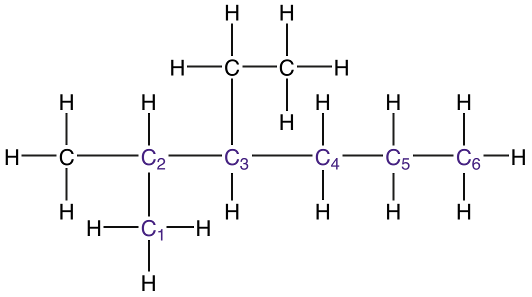
$$H_3C-\underset{\displaystyle CH_3}{\underset{|}{CH}}-\overset{\displaystyle CH_2CH_3}{\overset{|}{CH}}-CH_2-CH_2-CH_3$$

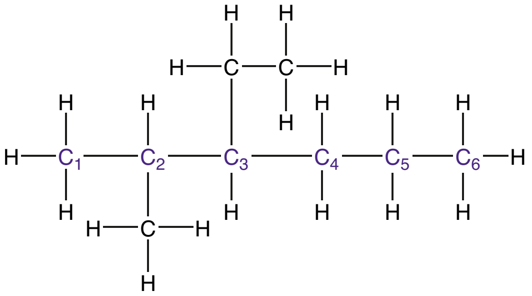
1. Identify the longest unbroken carbon chain involving the key functional group and give this the corresponding parent-name prefix.

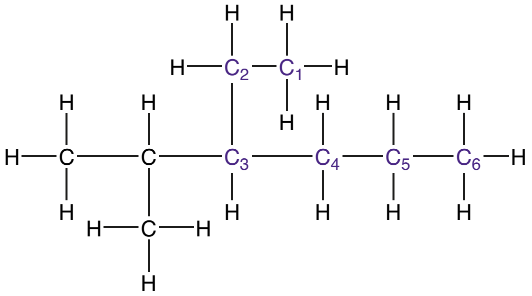






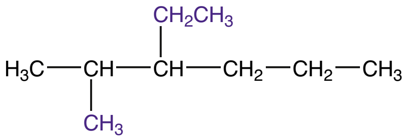






The longest unbroken carbon chain is 6 carbons in length and the molecule is saturated so is an alkane.

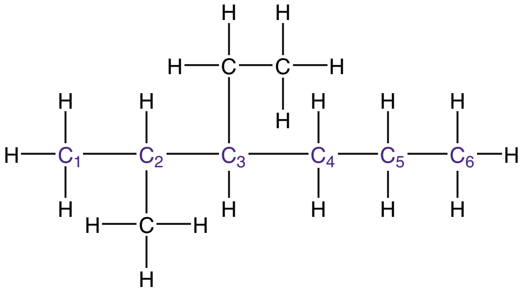
2. Locate and name the substituent group(s). When there are multiple different substituent groups, list them in alphabetical order, e.g. ethyl will come before methyl.

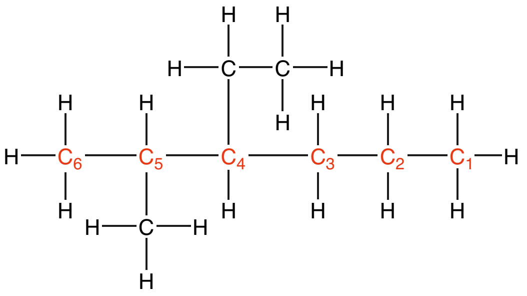


$$H_3C-\underset{\displaystyle CH_3}{\underset{|}{CH}}-\overset{\displaystyle CH_2CH_3}{\overset{|}{CH}}-CH_2-CH_2-CH_3$$

This molecule is a branched hydrocarbon where there is a methyl and an ethyl group.

3. Number each carbon atom in the parent hydrocarbon chain in a manner which results in the substituent groups being located at the lowest-numbered position on the hydrocarbon molecule.

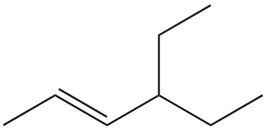




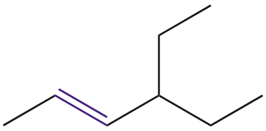
The lowest numbered positions for the substituent groups are with the methyl group on the second carbon and the ethyl group on the third carbon. Therefore, this molecule is named **3-ethyl-2-methylhexane.**

Naming a hydrocarbon drawn as a skeletal structure

Consider the following molecule:

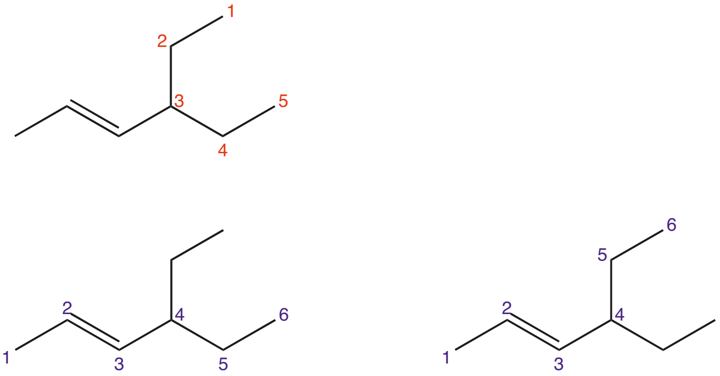


1. Locate and name the functional group(s). In the case of a skeletal structure remember that at each intersection of two lines this represents a carbon atom.



This hydrocarbon is an alkene as indicated by the C-C double bond.

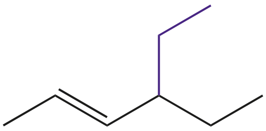
2. Identify the longest unbroken carbon chain involving the key functional group and give this the corresponding parent-name prefix.



Please note: the first option drawn above is not possible, as the functional group is not part of the main parent chain.

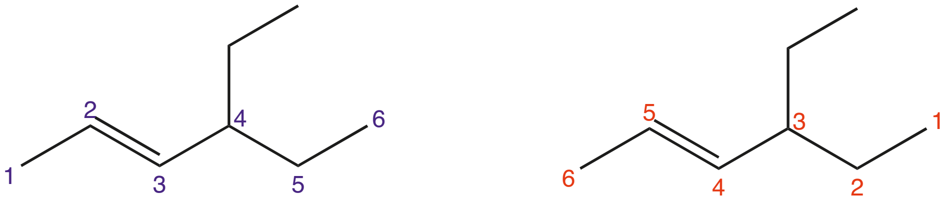
Hex due to the longest unbroken carbon chain containing the alkene group being 6 carbons.

3. Locate and name the substituent group(s).



This hydrocarbon is a branched hydrocarbon where the branch is an ethyl group.

4. Number each carbon atom in the parent hydrocarbon chain in a manner which results in the functional group being located at the lowest-numbered position on the hydrocarbon molecule.



The lowest numbered position for the double bond is on the second carbon. This results in the ethyl group being located on the fourth carbon. Therefore, the molecule is named **4-ethylhex-2-ene**.

Drawing Structural Formulae

The conventions for drawing structural formulae from a name are as follows:

1. Draw the main chain which is found in the name as meth, eth, prop, etc.
2. Draw all functional groups (if any).
3. Draw any substituent groups.
4. Fill in the rest of the molecule with hydrogen atoms ensuring all carbon atoms are sharing four pairs of electrons (they have four covalent bonds).

Drawing the structural formula for a hydrocarbon

Draw the structural formula for 3-methylpent-2-ene

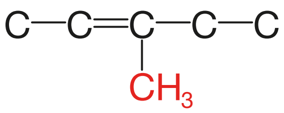
1. Drawing of just the pent- chain

$$C-C-C-C-C$$

2. Add in the double bond

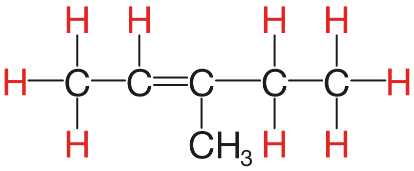
$$C-C=C-C-C$$

3. Add in the methyl group



$$C-C=\underset{CH_3}{\underset{|}{C}}-C-C$$

4. Fill in the rest of the molecule with hydrogen atoms.



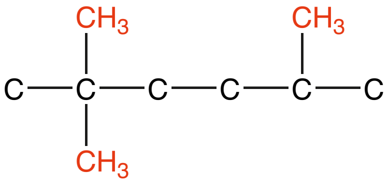
$$H-\overset{H}{\underset{H}{C}}-\overset{H}{C}=\underset{CH_3}{C}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

Draw the structural formula for 2,2,5-trimethyl-4-ethylhexane

1. Drawing of just the hex- chain

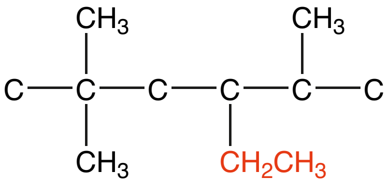
$$C-C-C-C-C-C$$

2. Add in the methyl groups onto the second and fifth carbon atoms



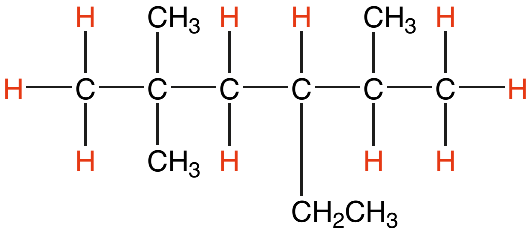
$$C-\overset{CH_3}{\underset{CH_3}{C}}-C-C-\overset{CH_3}{C}-C$$

3. Add in the ethyl group onto the fourth carbon atom



$$C-\overset{CH_3}{\underset{CH_3}{C}}-C-\underset{CH_2CH_3}{C}-\overset{CH_3}{C}-C$$

4. Fill in the rest of the molecule with hydrogen atoms.



$$H-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{CH_3}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{CH_2CH_3}{C}}-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

Questions

15. Given the following systematic names draw the corresponding structural formulae

    (a) octane

    (b) methylpropane

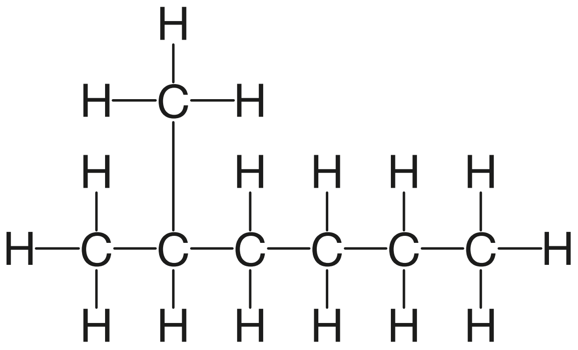
    (c) 3,4-diethylhexane

    (d) 6-ethyl-7-methyloct-3-ene

(8 marks) **KA4**

16. Name the following structural formulae

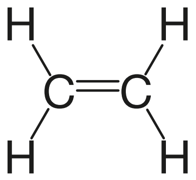
(a)



$$H-\overset{H}{\underset{H}{C}}-\overset{H-\overset{H}{C}-H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-H$$

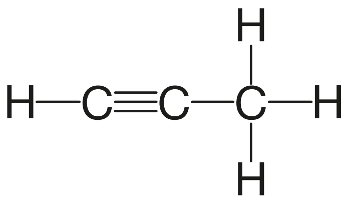
Name:

(b)



Name:

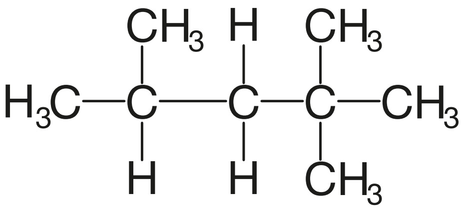
(c)



$$H-C\equiv C-\overset{H}{\underset{H}{C}}-H$$

Name:

(d)



$$H_3C-\overset{CH_3}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{CH_3}{\underset{CH_3}{C}}-CH_3$$

Name:

(8 marks) **KA4**

17. Draw the complete structural formula and state the systematic names of the four isomers of hexane.

Name ....

Name ....

Name ....

Name ....

(16 marks) **KA4**

Organic molecules and functional groups

Science understanding

Organic molecules have a hydrocarbon skeleton and can contain functional groups.

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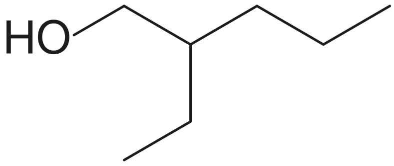
Organic molecules can also contain other functional groups involving different types of atoms. Functional groups are the atoms or groups of atoms that determine the chemical and physical properties of organic molecules.

Alcohols are a class or organic compounds that contain a functional group called the **alcohol** or **hydroxyl group** (R-**OH**), a **hydroxyl group**, where **R** relates to any hydrocarbon chain. Alcohols are useful in many different applications of organic chemistry not only the synthesis and production of commercial alcohol products. Alcohols are also used in the production of perfumes when reacted under specific conditions with acids and organic molecules containing acidic groups.

When naming an organic molecule with this functional group, the same rules apply as naming alkane, alkene and alkyne. Remember though that the parent chain needs to be the carbon chain containing the functional group, in this case the hydroxyl group. The suffix given to alcohol group containing organic molecules is **–ol**.

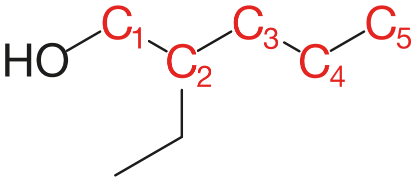
Example

Name the following molecule given the structural formula below:



1. **pent**

   The longest carbon chain containing the hydroxyl group is 5 carbons long.

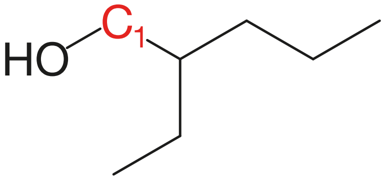


2. **pentan** (alkane)

   There are no double bonds within this structure

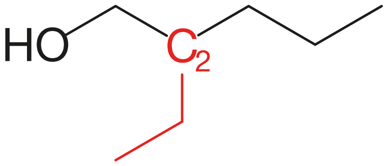
3. **pentan-1-ol**

   The hydroxyl group is located on the first carbon in the chain, as the numbering starts from the carbon that gives the functional group containing carbon the lowest number.



4. **2-ethylpentan-1-ol**

   An ethyl group is located on the 2nd carbon in the chain.



Other functional groups will be covered in the Stage 2 curriculum; below is a table of the functional groups you will encounter by the end of Stage 2.

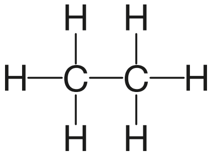
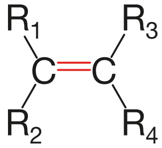
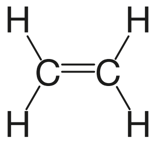
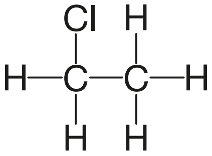
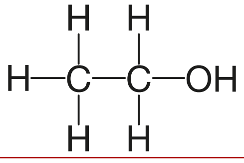
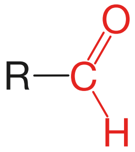
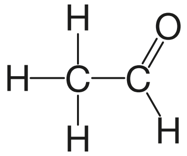
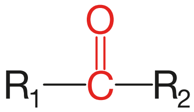
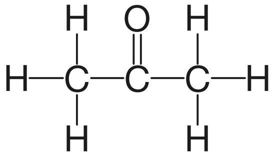
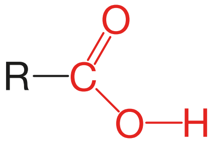
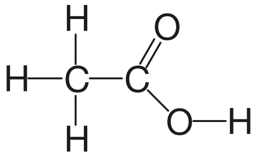
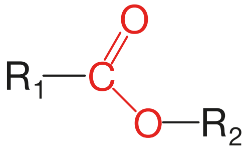
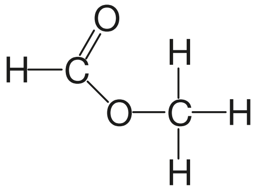
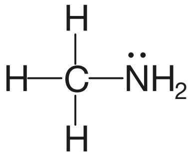
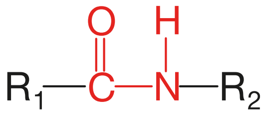
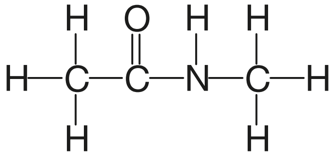
Homologous series	Functional group	Example
alkane	$R-H$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-H$
alkene		
alkyne	$R_1-C\equiv C-R_2$	$H-C\equiv C-H$
haloalkane	$R-X$	 $H-\underset{H}{\overset{Cl}{C}}-\underset{H}{\overset{H}{C}}-H$
alcohol	$R-OH$	 $H-\underset{H}{\overset{H}{C}}-\underset{H}{\overset{H}{C}}-OH$
aldehyde		
ketone	 $R_1-\overset{O}{\overset{\|}{C}}-R_2$	 $H-\underset{H}{\overset{H}{C}}-\overset{O}{\overset{\|}{C}}-\underset{H}{\overset{H}{C}}-H$
carboxylic acid		
ester		
amine	$R-\ddot{N}H_2$	 $H-\underset{H}{\overset{H}{C}}-\ddot{N}H_2$
amide	 $R_1-\overset{O}{\overset{\|}{C}}-\overset{H}{\overset{\mid}{N}}-R_2$	 $H-\underset{H}{\overset{H}{C}}-\overset{O}{\overset{\|}{C}}-\overset{H}{\overset{\mid}{N}}-\underset{H}{\overset{H}{C}}-H$

Table 3.07 Functional groups

Questions

18. Given the following systematic names, draw the corresponding structural formulae

    (a) ethanol

    (b) 2-methylpropan-2-ol

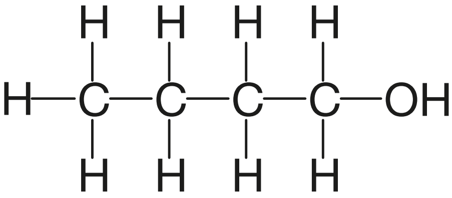
    (c) 4,4-diethyloctan-2-ol

    (d) 3-ethyl-4-methylhexan-1-ol

(8 marks) **KA4**

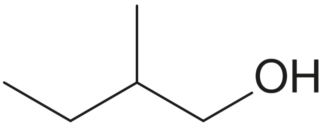
19. Systematically name the following molecules.

(a)



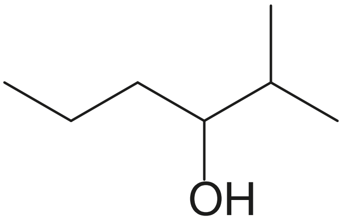
Name:

(b)



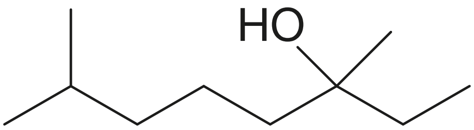
Name:

(c)



Name:

(d)



Name:

(8 marks) **KA4**

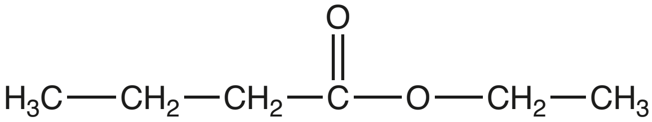
20. Esters are chemical compounds made by chemically combining an alcohol and a carboxylic acid. Esters have distinctive fruity aromas and one such ester has a pineapple aroma.

(a) Ethanol is used in the production of this ester. Draw the structural formula for this alcohol.

(2 marks) **KA4**

(b) The pineapple ester has the systematic name ethyl butanoate.

The structural formula for ethyl butanoate is given below.



$$H_3C-CH_2-CH_2-\overset{\overset{\displaystyle O}{\|}}{C}-O-CH_2-CH_3$$

Ethyl butanoate has a boiling point of 121°C whilst ethanol has a boiling point of 78°C.

(i) Explain why ethanol has a lower boiling point than ethyl butanoate.

(4 marks) **KA1**

(ii) Describe and explain how distillation separates the components of a mixture containing ethanol and ethyl butanoate.

(2 marks) **IAE4**

3.4: Polymers

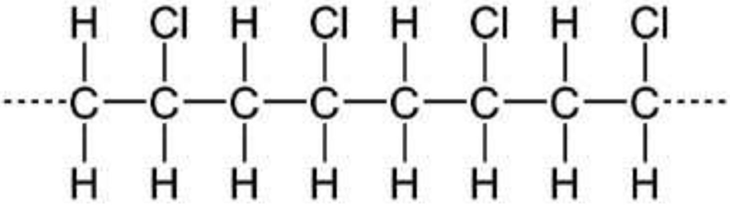
Introduction to polymers

Science understanding

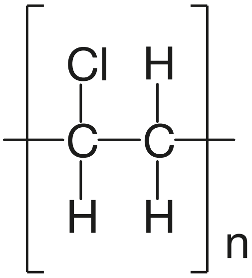
Polymers or macromolecules are very large molecules composed of small repeating structural units.

Identify the repeating unit of a polymer, given the structural formula of a section of a chain.

Polymers are very large long-chain molecules of high molecular mass consisting of repeating units called **monomers**. These monomers are repeated hundreds to thousands of times within one polymer molecule. The convention for identifying monomers within a polymer is to put square brackets around the repeating unit. Polymers can be represented by drawing the structural formula of a section of the polymer chain.



Polymers can also be represented by a formula where *n* indicates that there are *n* number of units of the repeating units.

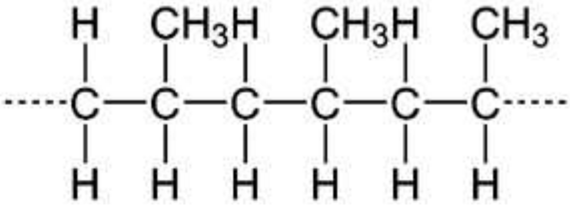


Many polymers are natural materials including cotton, starch, rubber and cellulose. Others are produced synthetically by materials scientists. The synthetic polymers include LDPE, HDPE and polyvinylchloride. Polymers make up many well-known materials that are used in everyday life ranging from PET bottles to car components.

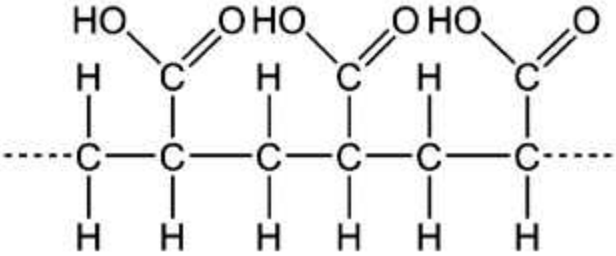
Questions

21. Identify the repeating unit using square brackets in the following segments of polymer chains.

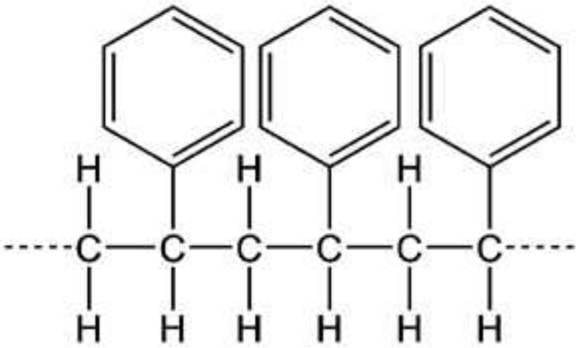
(a)



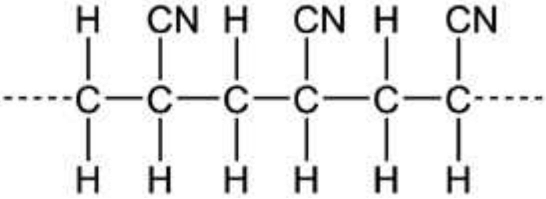
(b)



(c)



(d)



(8 marks) **KA4**

Addition polymers

Science understanding

Addition polymerisation occurs when monomer molecules link without the loss of atoms.

Addition polymers can be synthesised from alkene monomers.

Addition polymerisation is a chemical process whereby molecules (monomers) join together under specific conditions to create polymers. Each monomer used to make a polymer by addition polymerisation contains a double covalent bond between carbon atoms. By subjecting the monomers to heat and pressure while in the presence of a catalyst, the addition process shown below is able to be carried out.

Example

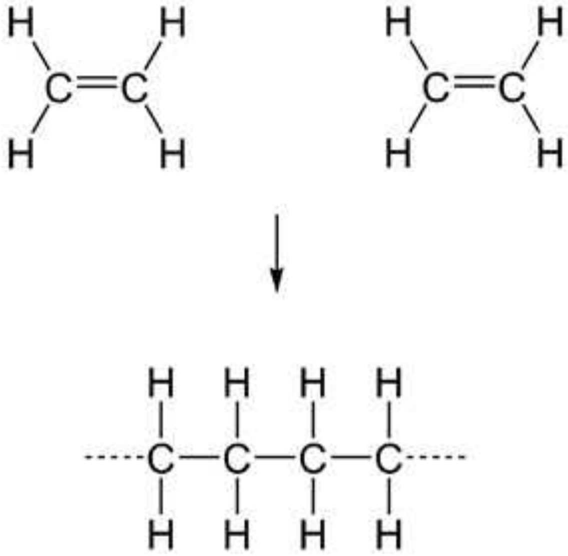


Figure 3.10 Polymerisation of ethene.

This process is used to make polyethylene which is the compound used to make plastic packaging.

The process carried out above is where a carbon-carbon double covalent bond within the original molecule (monomer) is replaced by a carbon-carbon single covalent bond. Therefore, the monomer is also able to be drawn from the polymer by first identifying the repeating unit and then replacing the carbon-carbon single covalent bond with a double covalent bond.

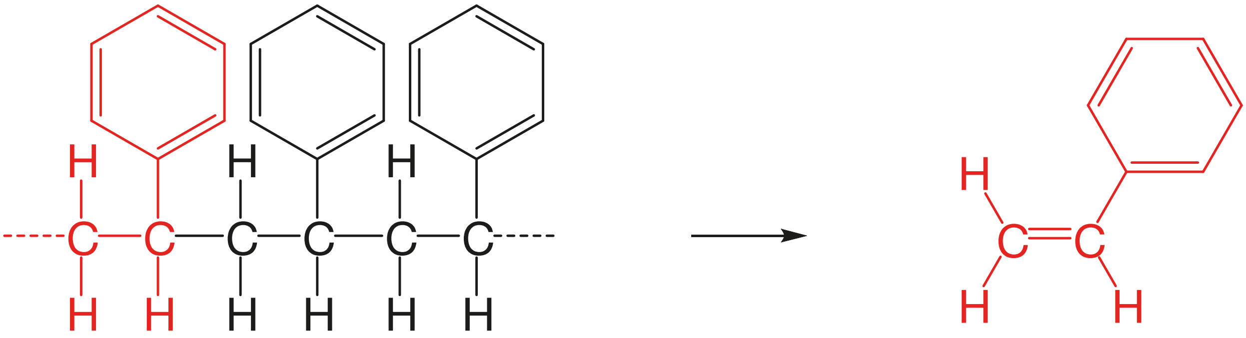


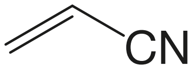
Figure 3.11 Structure of monomer from polymer.

Polystyrene is the compound used for producing disposable dinnerware and utensils as well in cooler boxes used to keep food fresh.

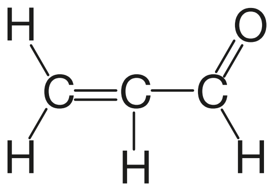
Questions

22. Using the monomers given below and the addition polymerisation process draw the structural formulae for the resulting polymers.

    (a) cyanoethene



    (b) propenal

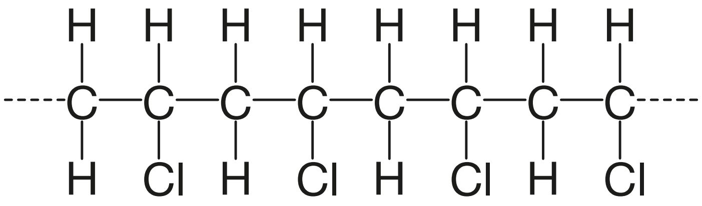


(4 marks) **KA4**

23. Each of the following polymers have been formed using the addition polymerisation process.

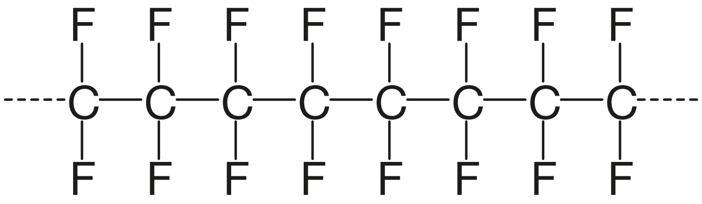
    (a) Identify the repeating unit in the following polymer structures using square brackets.

        (i) polyvinylchloride (PVC)



$$\begin{array}{ccccccccccccccccc}
 & H & & H & & H & & H & & H & & H & & H & & H & \\
 & | & & | & & | & & | & & | & & | & & | & & | & \\
\cdots & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & \cdots \\
 & | & & | & & | & & | & & | & & | & & | & & | & \\
 & H & & Cl & & H & & Cl & & H & & Cl & & H & & Cl &
\end{array}$$

        (ii) polytetrafluoroethene (Teflon)



$$\begin{array}{ccccccccccccccccc}
 & F & & F & & F & & F & & F & & F & & F & & F & \\
 & | & & | & & | & & | & & | & & | & & | & & | & \\
\cdots & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & \cdots \\
 & | & & | & & | & & | & & | & & | & & | & & | & \\
 & F & & F & & F & & F & & F & & F & & F & & F &
\end{array}$$

(4 marks) **KA4**

(b) Draw the structural formulae of the monomers from each of the previous polymers.

(i) polyvinylchloride

(ii) polytetrafluoroethene

(4 marks) **KA4**

Properties of polymers

Science understanding

Organic polymers have diverse properties and uses.

The properties of organic polymers depend on the interactions between the polymer chains.

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There are two main groups of polymers; **thermosetting polymers** and **thermoplastic** (softening) **polymers**.

Thermoplastic polymers are characterised by exhibiting weak intermolecular forces acting between the individual polymer chains within the material. The weak intermolecular forces are overcome by heating the material to its melting point.

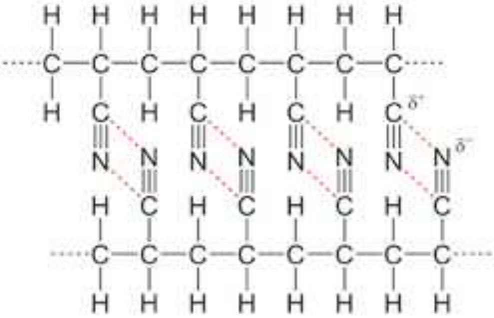


Figure 3.12: Cyanoacrylate (glue)

Thermoplastics are more suited to plastics where recycling is desired. When heated, thermoplastics soften and can then be reformed into the desired shape. Only a small amount of thermal energy is required to overcome the secondary interactions between chains and therefore separate those chains.

Thermosetting polymers are usually found in applications where a stronger more durable product is required. Thermosetting polymers are characterised by the presence of covalent bonds (intramolecular forces) between the individual polymer chains within the material. The strength of the covalent bonds between polymer chains results in the material having a more rigid structure. More thermal energy must be added to overcome the strong covalent bonds between polymer chains and cause the chains to separate and form a liquid. Thermosetting polymers do not melt. The individual atoms of carbon within the compound react with atmospheric oxygen (combustion) to form carbon dioxide before the melting point of the compound is reached. This results in the material burning or charring instead of melting.

These polymers cannot be used for recycling applications. An example of thermosetting polymers is found by reacting natural rubber polymers with compounds of sulfur under specific reaction conditions to form vulcanised rubber.

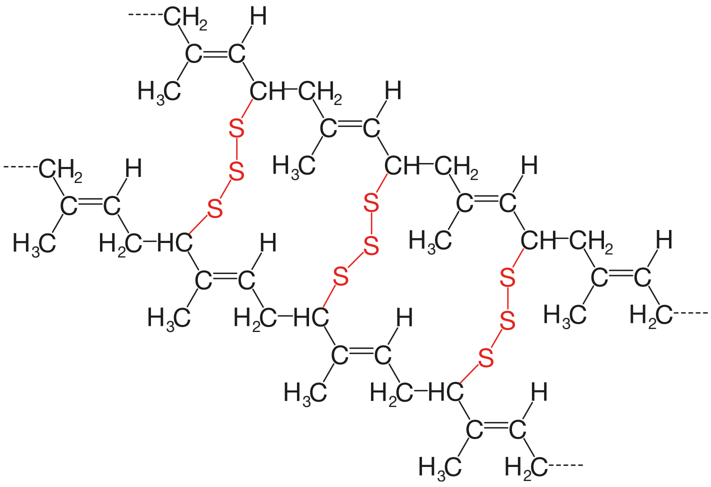


Figure 3.13 Vulcanised rubber

Science as a human endeavour:

Rotor blades for wind turbines are typically heavy, hard to transport and not recyclable due to them being produced using thermosetting polymers. Scientists and industry experts have collaborated to develop new rotor blades made from thermoplastic foam and composites that are lightweight and recyclable.

Helpful online resources:

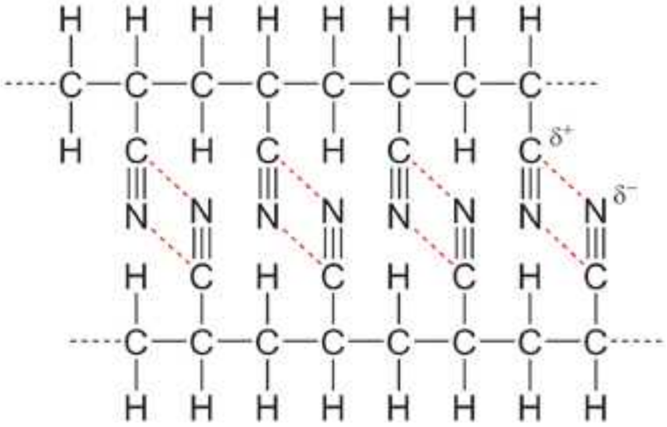
Lightweight rotor blades made from plastic foams for offshore wind turbines:

https://www.sciencedaily.com/releases/2016/10/161021084515.htm

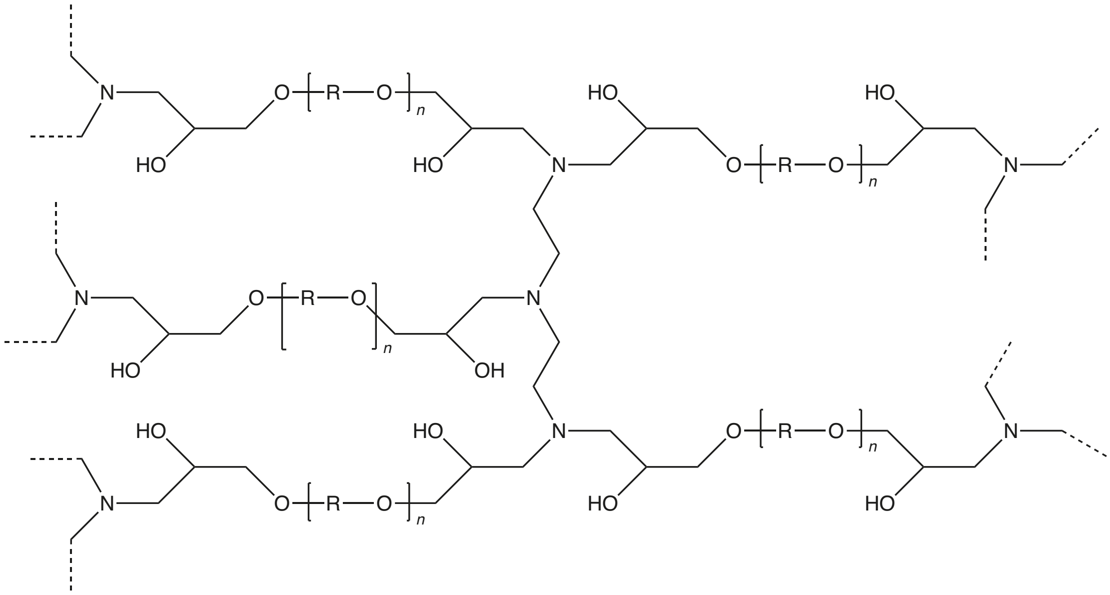
Question

24. Cyanoacrylate and epoxy resin are two adhesive products.

Their structural formulae are shown below.



cyanoacrylate



epoxy resin

Explain whether epoxy resin or cyanoacrylate would be more likely to withstand high temperatures with reference to their structures.

(4 marks) **KA2**

Summary Test 3: Molecules

1. Consider the following table of information:

Chloride	Melting point (°C)	Boiling point (°C)
silicon	–70	–57
sodium	808	1465

(a) State the chloride from the table above that is molecular.

.... *(1 mark)* **KA1**

(b) Name the particles present in the above compounds and state the name of the forces between the particles.

	Name of particles	Strength of forces
silicon chloride	....	....
sodium chloride	....	....

(4 marks) **KA1**

(c) Phosphorus trichloride has the formula PCl_3. Draw a diagram for this molecule showing covalent bonds and non-bonding electrons.

(2 marks) **KA4**

2. Sulfur and oxygen chemically combine to form sulfur dioxide.

(a) State the covalence for each of these elements.

(i) sulfur: ....

(ii) oxygen: ....

(2 marks) **KA1**

(b) Draw the electron dot structure for sulfur dioxide, SO_2.

(1 mark) **KA4**

(c) Sulfur is able to expand its octet. Explain what is meant by the term 'octet expansion' and why sulfur is able to do this.

....

....

....

....

....

....

....

.... *(4 marks)* **KA1**

(d) Draw a diagram to show the bonding and shape of a molecule of sulfur dioxide.

On this diagram show the polarity of the bonds using the appropriate conventions.

(3 marks) **KA1**

(e) Explain why the bond between S and O is polar.

(2 marks) **KA1**

(f) State and explain whether the sulfur dioxide molecule is polar.

(3 marks) **KA1**

(g) Name the secondary interaction present between sulfur dioxide molecules.

(1 mark) **KA1**

3. Nitrogen makes up approximately 80% of the Earth's atmosphere and is a relatively unreactive gas. However, in industrial applications it is reacted with hydrogen to produce ammonia.

(a) Draw the electron dot structure for nitrogen gas, N_2.

(1 mark) **KA4**

(b) Draw the structural formula for nitrogen gas and identify its shape.

(3 marks) **KA4**

(c) Explain why nitrogen is a relatively unreactive gas.

(2 marks) **KA1**

(d) Draw the electron dot structure for ammonia, NH_3.

(1 mark) **KA4**

(e) Draw the structural formula for ammonia and identify its shape.

(3 marks) **KA4**

(f) Explain why the ammonia molecule has this shape.

(2 marks) **KA1**

Gas	Molar mass (g mol^{-1})	Boiling point (°C)
nitrogen	28.02	–196
ammonia	17.034	–33

(g) Explain using diagrams why the boiling point of ammonia is much higher than that of nitrogen.

(5 marks) **KA1**

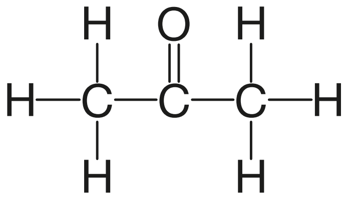
4. Propan-2-ol is an alcohol that can undergo oxidation in a laboratory to produce propanone, more commonly known as acetone. Acetone is the main compound found in nail polish removers.

   (a) Draw a structural isomer of propan-2-ol.

   (2 marks) **KA4**

   (b) When applied to fingernails, acetone acts as a solvent and dissolves any nail polish on the nail. Acetone also evaporates easily off of the nail due to its low boiling point.

   The structural formula of acetone is given below.



   Explain why acetone has a low boiling point.

   (2 marks) **KA2**

5. Polypropene is a polymer synthesised from an unsaturated organic compound (monomer) with the molecular formula C_3H_6 and systematic name propene.

   (a) Draw the structural formula for the monomer.

   (2 marks) **KA4**

   (b) Explain, with the aid of diagrams, whether you would expect the monomer propene or propan-1-ol to have a higher boiling point.

   (5 marks) **KA1**

(c) Upon completing the addition polymerisation process with the monomer, draw the structural formula for the polymer produced.

(2 marks) **KA4**

(d) Explain why polypropene has a much higher melting point than propene.

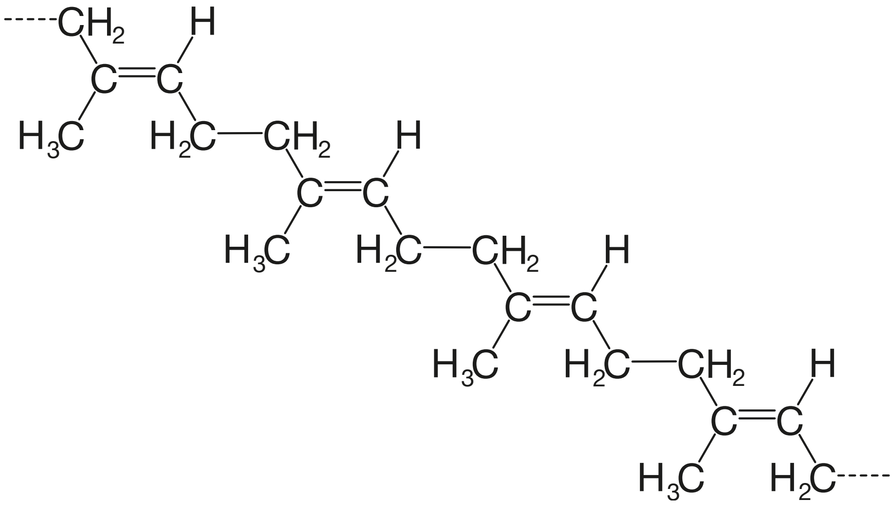
(4 marks) **KA1**

6. Polyisoprene is an organic compound that comes from the tree species *Hevea brasiliensis*. It is made up of many isoprene molecules.

(a) The systematic name of isoprene is 2-methylbuta-1,3-diene. Draw the structural formula for isoprene.

(2 marks) **KA4**

(b) The more common name for the polymer formed from this monomer is 'natural rubber', and its structural formula is shown below.



The properties of natural rubber may be improved by treating the rubber in a process known as **vulcanisation**.

Explain why vulcanisation improves rubber's ability to retain its shape at high temperatures.

(3 marks) **KA2**

Topic 4: Mixtures and solutions

Introduction

Mixtures are materials that contain two or more substances that have been mixed together but have not chemically bonded. Mixtures are found all around us: in a bottle of soft drink, in sea water and river water, and even in the air we breathe. Air is a mixture, as it is a combination of gases including nitrogen, N_2 (78%), oxygen, O_2 (21%), carbon dioxide, CO_2 (0.02%), and other gases such as argon (Ar), neon (Ne), helium (He), hydrogen (H_2), ozone (O_3), methane (CH_4) and water vapour (H_2O).

When a mixture's chemical composition is uniformly distributed, it is referred to as homogeneous. If the particle distribution is non-uniform, the mixture is termed heteregenous. A **solution** is a homogenous mixture of two or more substances. Solutions have two components: solute and solvent.

The **solute** is the substance that is dissolved, and the **solvent** is the substance in which the solute is dissolved. Sodium chloride dissolved in water can be used as an example of a solution. Here, solid sodium chloride is the solute which dissolves in water, which is the solvent.

When forming a miscible solution, or immiscible solution, the solvent is the substance that is present in the greatest amount.

All solutions have the properties that they possess due to the effect of three factors: molecular structure, pressure and temperature. These are the factors that affect a substance's solubility. **Solubility** refers to the amount of solute that will dissolve per unit of volume of solvent (at a particular temperature). This chapter, will focus predominantly on the effect of a substance's molecular structure upon its solubility.

Solutions are often classified as **concentrated** or **dilute** based on the relative number of particles in a given volume or, in other words, how much solute is in the solvent. However, in chemistry, this is defined more specifically. Solutions are described in terms of their:

- **Molarity (Molar concentration)** – the number of moles of solute per litre of solution (mol L^{-1})
- **Percentage mass** – the percentage of solute compared to solvent ($\frac{\text{mass of solute}}{\text{mass of solvent}} \times 100$)

4.1: Miscibility and solutions

Solubility and miscibility are related, yet distinct chemical terms. **Solubility** more generally describes a solute dissolving, mixing or combining within a solvent independent of the state of either of the solute or solvent. **Miscibility** describes liquids which can be mixed together to form a single liquid. The results can be considered as a solution of one liquid in another. A miscible solution is a homogeneous mixture that is formed when a solute is dissolved in a solvent.

In contrast, **immiscible** is the term used to describe liquids that do not mix. Two or more immiscible liquids form layers with interfaces between them. An immiscible solution is a heterogeneous mixture that is formed when a solute dissolves partially or not at all in a solvent. This concept is represented below in Figure 4.01.

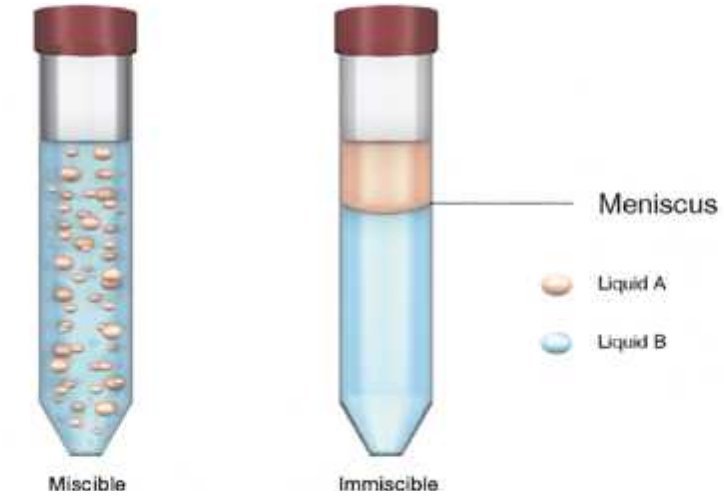


Figure 4.01: A miscible solution and an immiscible solution.

For a solute to dissolve in a solvent and form a solution, the particles in both liquids must be able to separate and mix together in a homogenous manner. The attraction among the particles in their original states must be replaced by attractions as strong as, or stronger than, those between solute and solvent particles. If the attraction between the solute and solvent particles are not as strong as (or stronger than) the solute, then the solute will not dissolve. This will result in an immiscible solution.

Polar and non-polar solvents

Science understanding

Solvents can be considered as polar (e.g. water, methanol) or non-polar (e.g. hexane, turpentine, petrol).

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Solvents can be classed as polar or non-polar. This classification arises from the combination and arrangement of atoms within the solvents. **Polar solvents** have significant dipole moments arising from polar bonds formed between atoms of differing electronegativities such as oxygen and hydrogen. Intermolecular forces of attraction form between the functional groups of polar solvents as shown in Figure 4.02.

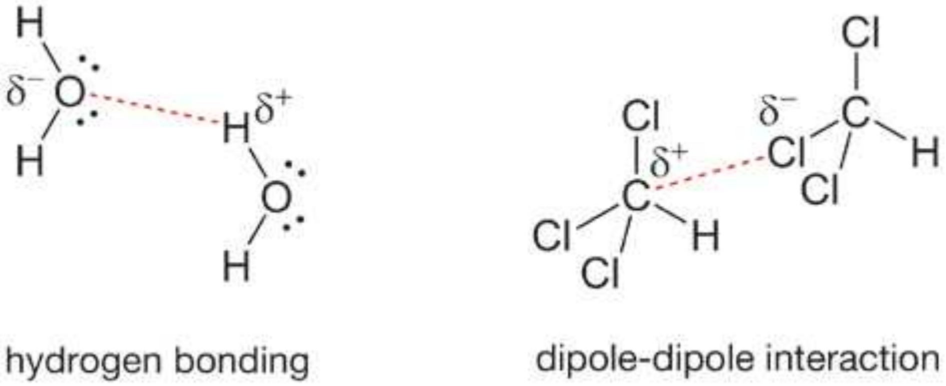


Figure 4.02: Examples of polar solvents and their intermolecular attractions between adjacent molecules, water (hydrogen bonds) and trichloromethane (dipole–dipole forces)

Non-polar solvents are composed of non-polar molecules. Non-polar bonds arise when covalent bonds form between atoms of similar electronegativities, and non-polar molecules arise from the molecule being deficient of a net dipole. The molecules present within a non-polar solvent interact via weak intermolecular forces of attraction. The strength of the intermolecular forces between non-polar molecules increases with molecular size. Cyclohexane as a non-polar solvent is shown below in Figure 4.03.

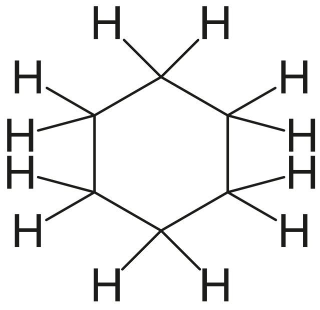


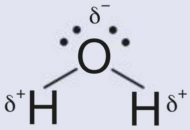
Figure 4.03: Cyclohexane – consists of non-polar bonds, and therefore is non-polar over all.

Science as a human endeavour

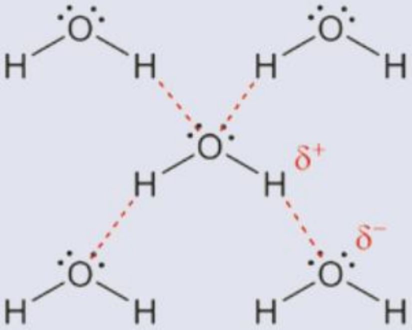
Water: the universal solvent

Water is the most commonly known, and referred to solvent. This “universal solvent” dissolves the widest range of substances, allowing aqueous solutions of water to contain a vast number of dissolved nutrients, minerals and compounds. It is one of the reasons why water is imperative for all life on Earth.

Water possesses this ability due to its molecular structure, as a covalently bonded, v-shaped molecule. As discussed in Chapter 3, water is classed as a polar molecule, as the centres of positive and negative electric charge are in different locations.



Water is referred to as a polar solvent as it's molecules are held together by hydrogen bonds, strong intermolecular forces of attraction that form between molecules.



Water molecules are electric dipoles, allowing water to be attracted to a vast number of different types of compounds. The attractive force exhibited by water, is often strong enough to disrupt a compounds interionic forces resulting in a compound dissolving. This is explored later in this Topic.

Solubility of polar and non-polar molecules

Science understanding

Highly polar molecular substances are more soluble in water than non-polar molecules of a similar size.

Polar and non-polar solvents do not readily mix.

The strengths of intermolecular interactions are important factors that govern whether a molecular substance dissolves in another. To determine this, their polarity must first be considered.

Non-polar solute and non-polar solvent

Non-polar substances are likely to dissolve in other non-polar substances. To examine this, iodine in cyclohexane can be considered. Both molecules are largely non-polar as they contain non-polar bonds, and therefore adjacent molecules are held together via weak dispersion forces.

In terms of the two liquids coming together as a homogenous solution, the molecules of cyclohexane and the molecules of iodine are separated from each other. Upon mixing, molecules of iodine and cyclohexane come together, as the intermolecular attraction is weak between the original molecules. This allows the subsequent formation of a homogenous solution as shown in Figure 4.04 below.

Figure 4.04: The solvent (red) and solute (blue) particle’s dispersion forces are separated, allowing the molecules to mix and new intermolecular forces (still dispersion forces) to form between the miscible solution of red and blue particles.

Polar solute and polar solvent

Polar solutes are likely to dissolve in polar solvents and form a miscible solution. For example, ethanol (polar solute) dissolves in water (polar solvent) forming a homogeneous solution. In their original states, both ethanol and water use hydrogen bonds to hold their adjacent molecules together.

Molecules of water interact strongly in solutions through the formation of intermolecular forces of attraction (hydrogen bonds). When a volume of ethanol is added, the molecules of ethanol disrupt the extensive hydrogen bonding between the molecules of water. This action causes the molecules of water to separate and form a homogenous solution with the molecules of ethanol. This is shown in Figure 4.05.

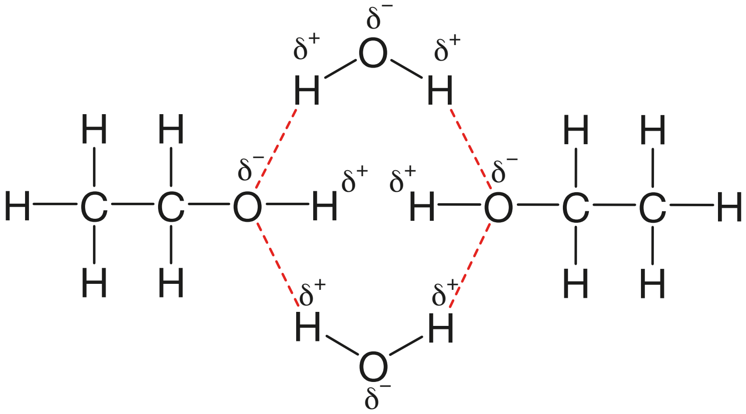


Figure 4.05: Miscible solution formed between molecules of ethanol and water. The solution is formed through hydrogen bonds between molecules of both substances.

Solvent and solute pair of opposing polarities

When the polarities of solutes and solvents are dissimilar, then, as a general rule, they do not mix, resulting in immiscible solutions. Polar solvents are immiscible in non-polar solvents.

To explore this, non-polar iodine (I_2) could be placed into polar water. Molecules of iodine are held together by weak dispersion forces, and water molecules are held together by stronger hydrogen bonds. Molecules of iodine cannot overcome the hydrogen bonds operating between water molecules and, hence, iodine is insoluble in water at lower temperatures. This results in the visible formation of two layers, and an immiscible solution.

However, certain non-polar materials including the hydrocarbons methane, ethane and propane will all partially dissolve in water and form a solution.

Table 4.01 summarises the miscibility of differing solute/solvent combinations based on their polarity.

Solute	Polar solvent	Non-polar solvent
Polar	Soluble	Insoluble
Non-polar	Insoluble	Soluble

Table 4.01: General miscibility of solute/solvent pairs based on polarity

Questions

1. Brass is an alloy, a solid mixture, which is manufactured from molten copper and zinc. Modern brass is 67% copper and 33% zinc.

   Describe how when molten, these metals represent a solute and solvent interaction.

....

....

.... *(2 marks)* **KA1**

2. The rubber industry uses gasoline as a non-polar solvent in the preparation of rubber cement, as well as quick-drying oil vanishes and dyes. It is a complex liquid mixture of over 150 hydrocarbons, with a variable composition dependent on the crude oils used, and the refinery processes available.

   (a) Using gasoline as an example, explain what is meant by the term miscible.

....

....

.... *(2 marks)* **KA1**

   (b) Upon mixing mixing gasoline with water, two layers form. State the type of solution now produced.

.... *(1 mark)* **KA1**

3. Ammonia is often described as a “water like” solvent.

   (a) Describe the intermolecular attractions that would be most significant between molecules of ammonia.

....

....

.... *(2 marks)* **KA1**

   (b) Explain how this is similar to how water behaves as a solvent.

....

....

.... *(2 marks)* **KA1**

   (c) Within each pair of compounds below, identify and explain the substance that is more likely to be miscible with ammonia. Explain your predictions.

      (i) Ethanol (C_2H_5OH) and cyclohexane (C_6H_{12})

....

....

....

.... *(3 marks)* **KA1**

      (ii) Iodine pentachloride (ICl_5) and propan-1-ol (C_3H_7OH)

....

....

....

.... *(3 marks)* **KA1**

4. It is observed that a substance dissolves in hexane but not in water.

   (a) Explain what polarity the substance is likely to exhibit, and explain the observation.

   (3 marks) **KA4**

   (b) Explain the colloquial statement ‘like dissolves like’ in chemical terms.

   (1 mark) **KA1**

5. Predict the miscibility of the following solute/solvent pairs. Explain your prediction.

   (a) methanol and benzene

   (3 marks) **KA1**

   (b) propan-1-ol and water

   (3 marks) **KA1**

   (c) octane and cyclohexane

   (3 marks) **KA1**

Solubility of small and large molecules with similar polarity

Science understanding

Molecular substances with small molecules are more soluble in water than larger molecules of similar polarity.

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The size of a molecule influences its miscibility in water. The larger the solute particles are, the more difficult it is for solvent particles to separate and surround the molecules. The molecules within the solvent must be able to disrupt the existing intermolecular forces operating between solute molecules. Once the solvent has separated the particles (molecules/atoms/ions) within the solute, then the molecules of the solvent will surround the solute particles.

This results in a general rule: as the size and subsequent molecular mass of a molecule increases, its miscibility decreases.

Example

Methanol (CH_3OH) and ethanol (CH_3CH_2OH) are miscible in water, however, hexan-1-ol ($CH_3CH_2CH_2CH_2CH_2CH_2OH$) is immiscible even though it is capable of some hydrogen bonds with water.

This is predominantly due to the intermolecular forces that exist between molecules of hexan-1-ol. Hexan-1-ol has six carbon atoms per molecule and one polar hydroxyl group.

Dispersion forces operate between carbon atoms on adjacent molecules, and hydrogen bonds operate between hydroxyl groups. Larger organic molecules have more carbon atoms and, hence, a greater number of sites over which dispersion forces will act. Dispersion forces attract adjacent molecules of hexan-1-ol together. Molecules of hexan-1-ol come very close together. The magnitude of an attractive (electrical) force increases as the molecules come closer together. Stronger forces require more energy to separate.

Molecules within the solvent are simply unable to separate the molecules of hexan-1-ol given the strength of the attractive forces operating between molecules. As no separation occurs, an immiscible solution results.

Science as a human endeavour

Fat-soluble and water-soluble vitamins

Vitamin C and vitamin D3: vitamins C and D3 are both essential for the maintenance of positive human health. Vitamin C has an important role as an antioxidant, and is involved in the maintenance of the immune system. Vitamin D3 is required for the absorption of many minerals, including calcium, and therefore is essential for good bone health. The structures of both vitamins are shown in Figure 4.06

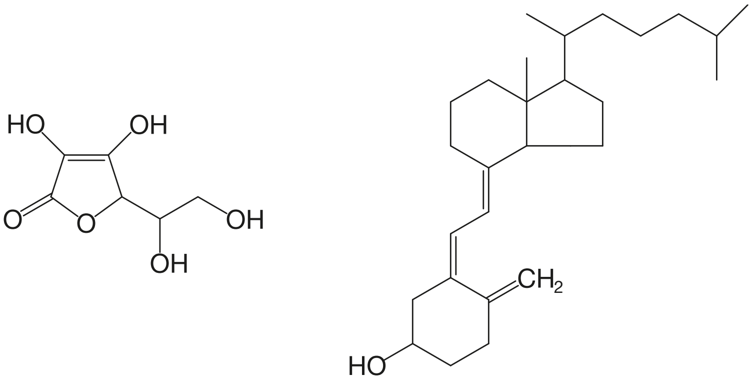


Figure 4.06: Structures of vitamin C and vitamin D3

If these structures are compared, vitamin C has four hydroxyl groups and a carbonyl group, all of which are capable of hydrogen bonding with water. It has a relatively small non-polar hydrocarbon component and, due to this, is highly miscible in water. Vitamin D3 has a larger molecular mass and contains only one polar group. There are many sites over which dispersion forces act between molecules of vitamin D3. Molecules of water cannot overcome the extensive intermolecular forces operating between molecules of vitamin D3.

The overall effect of this is that vitamin C is considered a water-soluble vitamin, and vitamin D is considered a fat-soluble vitamin (as fat is considerably non-polar). The solubility of vitamins, other than vitamin C and vitamin D3, is summarised below in Table 4.02.

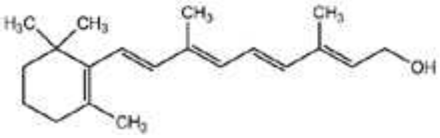
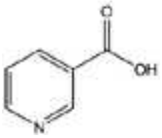
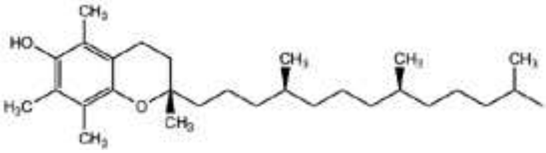
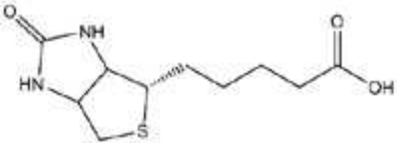
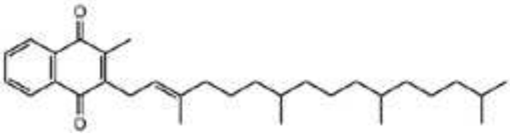
Vitamin	Water-soluble	Fat-soluble	Structure
Vitamin A (Retinol)		✓	
Vitamin B1 (Thiamine) Vitamin B2 (Riboflavin) Vitamin B3 (Niacin) Vitamin B5 (Pantothenic Acid) Vitamin B6 (Pyridoxine) Vitamin B9 (Folic Acid) Vitamin B12 (Cobalamin)	✓		 Vitamin B3 – an example of a B-group vitamin
Vitamin E (Tocopherols and Tocotrienols)		✓	
Vitamin H (Biotin)	✓		
Vitamin K		✓	

Table 4.02: Solubility of the main vitamins required for good human health

Questions

6. Water is a polar solvent. Methanol is a small organic molecule that is miscible in water. Consider this information when answering the following:

   (a) Draw a diagram to show the hydrogen bonds that form between water molecules and methanol molecules when mixed.

   (2 marks) **KA4**

   (b) Methanol is also capable of forming a homogenous solution when mixed with hexane. Explain this in terms of its structure.

   (2 marks) **KA4**

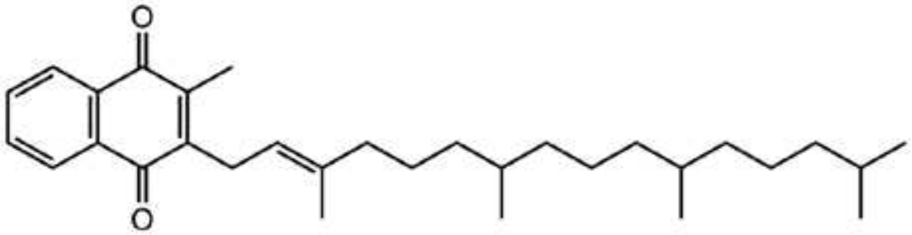
(c) Predict heptan-1-ol's solubility in water. Explain your prediction.

(2 marks) **KA1**

(d) Explain the general requirement for an organic compound to be soluble in water.

(2 marks) **KA1**

7. Vitamin K plays an important role in blood clotting. Consider its structure below.



(a) Describe what would occur in terms of intermolecular attractions when when vitamin K is mixed with water.

(3 marks) **KA1**

(b) Consequently, state if Vitamin K is considered a fat-soluble or water-soluble vitamin?

(1 mark) **KA1**

Mixing polar and non-polar substances

Science understanding

Compounds with non-polar and polar or ionic components facilitate the mixing of polar and non-polar substances

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Due to their opposing polarities, generally, polar and non-polar substances do not mix. Substances that have both non-polar and polar, or ionic components are useful for making two immiscible substances mix together.

Example

Oil and water

Two common immiscible substances are cooking oil and tap water, oil being predominantly non-polar, and tap water significantly polar. This is why water alone cannot be used to remove oil from a greasy pan, and why oil and water cannot be mixed together to form a solution. The molecules within cooking oil have a lower density than an equivalent volume of water. A third substance, called an **emulsifier**, can facilitate the mixing of the two immiscible liquids. The substance most commonly used is **detergent**. The structure of an ion of detergent is shown below in Figure 4.07.

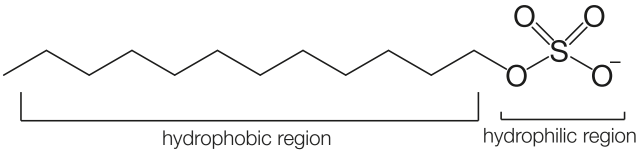


Figure 4.07: Structure of a detergent ion

Detergent has a chemical structure with a "head" and a "tail". The tail is a larger, non-polar hydrocarbon that does not mix well with water (it is **hydrophobic**). The head is a polar ionic sulfate ($R\text{-}SO_4^-$) or sulfonate ($R\text{-}SO_3^-$) that is miscible with water (it is **hydrophilic**). The head is polar and ionic, as one oxygen is attached via a single bond to the sulfur atom and a remaining unbonded electron. This results in an overall negative charge. The charge in the oxygen atom is a full negative charge. A full charge is greater in magnitude than a partial charge. The magnitude of an attractive force increases with charge. Intermolecular forces between a partial charge (of a dipole system) and an ion are termed ion–dipole forces. This is explored in more detail in subtopic 4.2.

Detergent ions are attracted to both water and oil. When detergent is added, the polar head on the detergent ion is attracted to a water molecule and the detergent tail is attracted to an oil molecule. The detergent creates a mixture of water with droplets of oil spread through it. The result is termed an **emulsion**, whereby the dispersion of oil into water has occurred. This concept is depicted below in Figure 4.08.

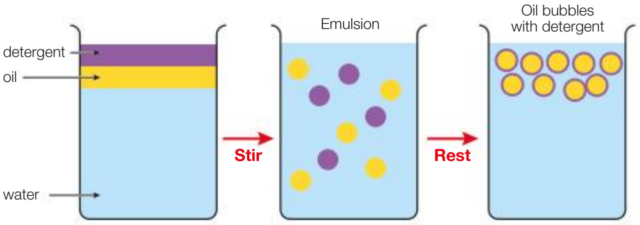


Figure 4.08: The emulsion that is created when detergent is added to the immiscible solution of oil and water.

An emulsifier is added to attain stability of the emulsion. In this example, the detergent is the emulsifier. The emulsifying action of detergent allows the oil to remain dispersed through the water. Specifically, the detergent ions do this by forming a **micelle**. A micelle is a spherical structure in which the non-polar tails of the detergent ions (which are attracted to the oil droplet), move to the centre of the molecule. The non-polar tails dissolve in the mass of grease, whereas the polar heads dissolve in water. This is important, as the heads are negatively charged, which prevents the aggregation of multiple micelles together, leading to the oil remaining dispersed through the water. Micelle formation is crucial to the function of detergents and soaps. An example of a micelle is depicted in Figure 4.09.

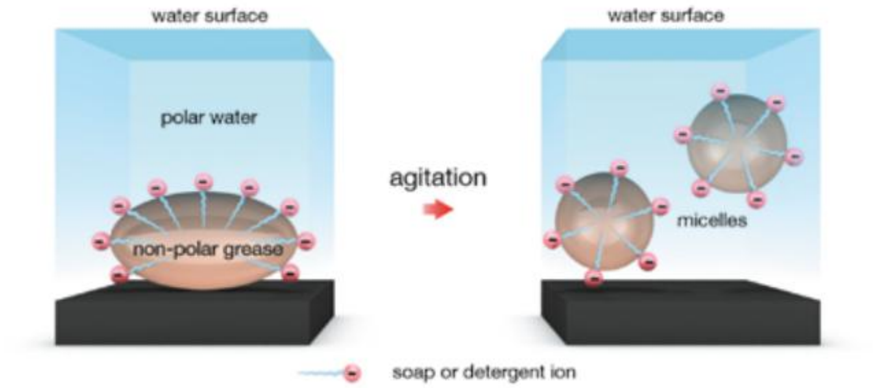


Figure 4.09: Removal of grease by micelle formation

Science as a human endeavour

Chemistry in cooking: mayonnaise

Two of mayonnaise's main ingredients, vinegar and oil, do not mix. Vinegar consists of polar molecules, namely water and ethanoic acid. Hydrogen bonds hold adjacent acetic acid and water molecules together. When combined, the water and ethanoic acid molecules are miscible resulting in the vinegar solution (a homogenous solution). Oil molecules have a large molecular size (12–22 carbon atoms per molecule) and there are many sites over which dispersion forces will act. Therefore molecules of ethanoic acid cannot overcome the extensive (dispersion) forces of attraction which operate between the oil molecules. Since molecules of ethanoic acid cannot (1) separate and (2) surround the molecules of oil, ethanoic acid is not soluble in oil. This would result in an unappealing mayonnaise that does not have the desired taste. To allow these two substances to mix, egg yolks are used as an emulsifier. The lecithin within egg yolks is the specific emulsifying agent. Lecithin is shown in Figure 4.10.

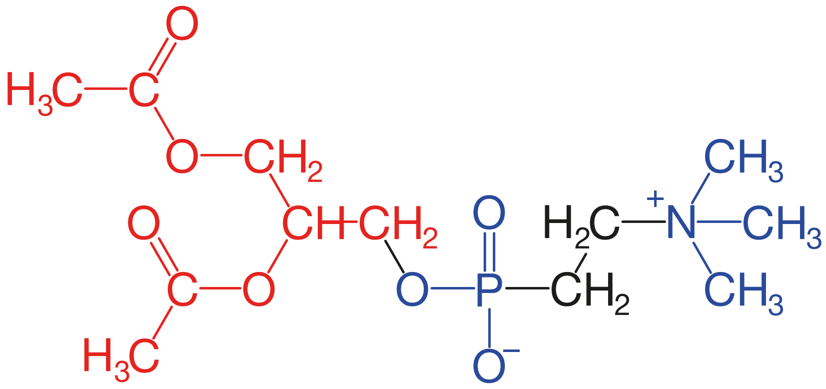


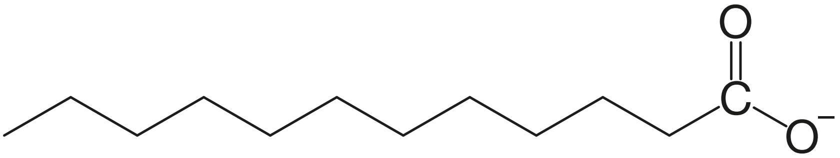
Figure 4.10: Structure of lecithin found within egg yolks.

As shown above, lecithin has a non-polar component (red) which is able to form intermolecular interactions with oil, and an ionic component (blue) which forms strong intermolecular attractions with water. This process of the lecithin allowing vinegar and water to be dispersed among one another forms the emulsion that is known as mayonnaise.

Questions

8. A molecule of soap is shown below.

   (a) Label its structure with the following terms: hydrophobic, polar, hydrophilic and non-polar.



(2 marks) **KA4**

   (b) With reference to the structure above:

      (i) Explain what is meant by the terms hydrophobic and hydrophilic.

....

....

.... *(2 marks)* **KA1**

      (ii) Detergents differ from soaps as they can have a sulfonate head, instead of the carboxyl head possessed by soap molecules. Describe how a detergent facilitates the mixing of a polar and non-polar substance.

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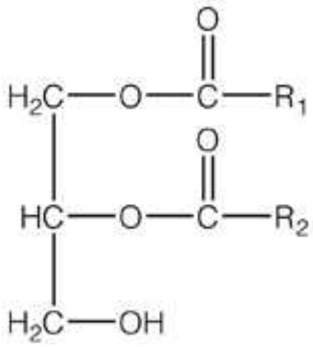
.... *(3 marks)* **KA1**

9. Ice cream is prepared from milk and cream,making it a mixture of fat and water. Diglycerides are often added to commercially produced ice creams to allow them to have a creamy texture. The general structure of a diglyceride is shown below.



(a) Explain how the diglyceride molecule acts as an emulsifier in the ice cream mixture.

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.... *(3 marks)* **KA1**

(b) Describe how an emulsion differs from an emulsifier.

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.... *(2 marks)* **KA1**

4.2: Solutions of ionic substances

An ionic compound is one which consists of ions held together by electrostatic forces. More specifically, these ions are a positively charged ion (**cation**) and a negatively charged ion (**anion**). Ionic compounds form continuous lattices and therefore are crystalline solid salts.

When placed in an aqueous solvent (most commonly water), an interaction immediately begins to take place between the polar water molecules and the anions and cations in the salt. Water molecules are electric dipoles and contain partial positive and negative poles. The anions (–) and cations (+) within the lattice of the solid salt carry a full magnitude of electric charge. Electrical forces of attraction operate between the anions and cations within the lattice and the molecules of water (dipole). The forces of attraction are termed **ion–dipole** interactions. The hydrogen atoms of the water molecules (positive pole) are attracted to the anions (negative pole) within the lattice. The oxygen atom of the water molecules (negative pole) is attracted to the cations (positive pole) within the lattice. The molecules of water become bound to the anions and cations at the surface of the ionic lattice. If the force of attraction between the ions and the molecules of water is greater in magnitude than the ionic bonds within the lattice, then the ions become **dissociated** from the lattice. Molecules of water separate anions and cations from the lattice and then surround the ions, forming a **hydration shell**. The hydration shell prevents the ions from recrystallising.

If the attraction is stronger between the water molecule and ion, in comparison to the interionic (ion–ion) attraction, then the salt is soluble. If the ion–ion attraction is stronger than the water–ion attraction, the salt is insoluble in water.

On an atomic level, the ions in the salt are hydrated by the water molecules. Hydration is the process of attraction and association of molecules of a solvent with molecules or ions of a solute. As ions dissolve in a solvent they spread out and become surrounded by solvent molecules. This is depicted in Figure 4.11 for the hydration of sodium chloride.

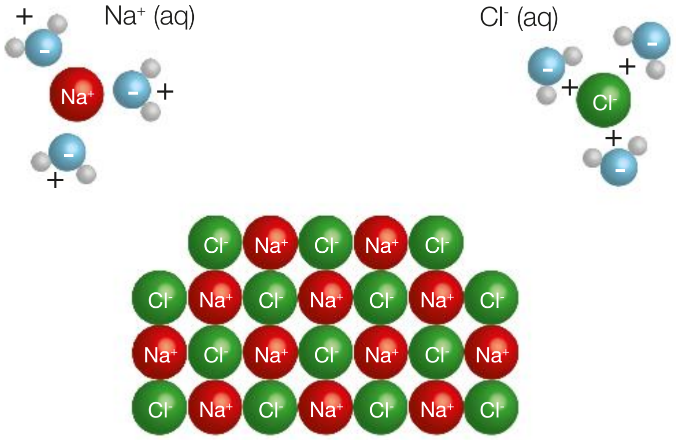


Figure 4.11: Hydration of Na^+ and Cl^- ions by water.

It is important to consider the orientation of the water molecules to the positive sodium ions and negative chloride ions. Figure 4.12 shows the depiction of hydrated ions:

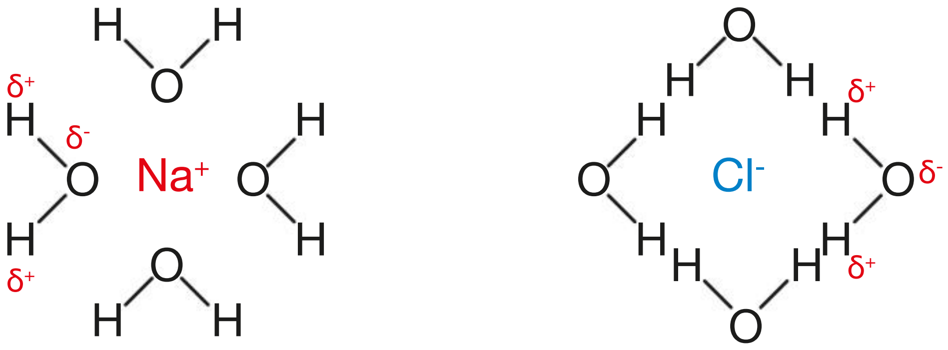


Figure 4.12: Hydrated sodium cations and chloride anions.

Note that the oxygen atoms (δ–) of the water molecules are attracted to the dissociated cation (Na^+) and the hydrogen atoms (δ+) are attracted to the dissociated anion (Cl^-).

Science as a human endeavour

Road Salt

In parts of the world where extreme cold and snow are experienced, driving on roads can become particularly hazardous. This hazard arises from the formation of ice, which reduces the friction experienced between the tyres of a vehicle, and the road's surface. Road salt, is commonly applied to halt the formation of this problematic ice. It is made from halite, which is the natural, mined mineral of sodium chloride.

Chemically, road salt works by inducing freezing point depression. If bad weather is predicted, the road salt is applied to a damp road where it dissociates into its component ions. The resulting salt solution, has a lower freezing point, making it harder for the water on the road to solidify. When applied to an already icy road, the same principle applies. Here, the solution spreads under the ice, loosening it's bond with the road's surface. This then allows mechanical means of removal to be more effective.

The solubility of salts is dependent on their composition and, like many concepts in chemistry, this solubility follows general rules, as shown in Table 4.03.

Type of anion	Formula	Solubility of common salts
Acetates	CH_3COO^-	All soluble
Nitrates	NO_3^-	All soluble
Chlorides	Cl^-	All soluble except Ag^+, Hg^+ and Pb^{2+}
Sulfates	SO_4^{2-}	All soluble except Ba^{2+}, Sr^{2+} and Pb^{2+} (Ca^{2+}, Ag^+ and Hg^+ are slightly soluble)
Carbonates	CO_3^{2-}	All insoluble except Na^+, K^+ and NH_4^+
Hydroxides	OH^-	All insoluble except Na^+, K^+ and NH_4^+
Sulfides	S^{2-}	All insoluble except Na^+, K^+ and NH_4^+ and the Group II ions Mg^{2+}, Ca^{2+}, Ba^{2+} and Sr^{2+}
Type of cation	**Formula**	**Solubility of common salts**
Sodium	Na^+	All soluble
Potassium	K^+	All soluble
Ammonium	NH_4^+	All soluble

Table 4.03: Solubility of common salts

Questions

10. Calcium chloride ($CaCl_2$) is another salt commonly used to remove ice from roads. It resembles sodium chloride (NaCl) as it is a white, crystalline solid.

    (a) The salt is made from the reaction between calcium hydroxide and hydrochloric acid.

    Explain how the salt is formed referring specifically to the forces that hold the ions together.

....

....

.... *(2 marks)* **KA1**

    (b) Calcium chloride is soluble in water.

    Describe what occurs at the atomic level when the salt comes into contact with molecules of water.

....

....

....

.... *(3 marks)* **KA1**

    (c) Draw and label a diagram of a sodium chloride lattice to depict what is meant by the following terms:

    (i) hydration (ii) dissociation

(4 marks) **KA4**

    (d) All chloride salts are soluble apart from silver, mercury and lead chloride.

    Explain in terms of interionic attraction why these salts are insoluble.

....

....

.... *(1 mark)* **KA1**

11. Complete the following table by putting a tick in each relevant column:

Salt	Soluble in water	Insoluble in water
iron(III) nitrate		
copper(II) sulfate		
mercury(II) sulfide		
zinc chloride		
ammonium phosphate		
lead(II) acetate		

(6 marks) **KA4**

Ion-dipole interactions and ionic equations

Science understanding

Many ionic substances are soluble in water. This is particularly so for ammonium and alkali metal salts.

Equations can be written to represent the dissociation and hydration of ions that occurs when ionic substances dissolve in water.

When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, the ions are physically dissociated from the lattice.

Science as a human endeavour

Svante August Arrhenius – ionic dissociation

Ionic dissociation was first suggested by a Swedish chemist by the name of Svante August Arrhenius (1859–1927) (Figure 4.13) as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry.

Figure 4.13: Svante August Arrhenius

Consider table salt (NaCl), a chloride salt containing anions and cations that are highly soluble in water. A chemical equation is used to represent this process. This is a complete ionic equation in which the dissolved ionic compound(s) are written as separated ions.

Note: insoluble or slightly soluble salts dissociate when placed in water, however only to a slight partial degree. Their dissociation is represented with reversible arrows ($\rightleftharpoons$).

As shown on page 174, the sodium and chloride ions separate via dissociation. This dissociation can be summarised by the following equation.

$$NaCl_{(s)} + aq \rightleftharpoons Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

The resultant solution is called an **electrolyte**. Sodium chloride is soluble, as the ion–water attraction is stronger than the ionic bonds between adjacent anions and cations within the lattice. This ion–water attraction in chemistry is more correctly described as the ion–dipole attraction, and the subsequent formation of an **ion–dipole interaction**.

An ion–dipole interaction is an attractive force that results from the electric attraction between an ion and a neutral molecule that has a dipole.

Ion–dipole interactions are most commonly found in solutions and are especially important for solutions of ionic compounds in polar liquids. In the above example, both the sodium and chloride are ions, and the neutral molecule with a dipole is water. The sodium cation attracts the partially negative end of the polar water molecule, as shown in Figure 4.14.

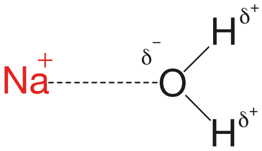


Figure 4.14: Cation–dipole interaction.

The chloride anion attracts the partially positive end of the polar water molecule, as shown in Figure 4.15.

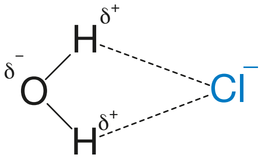


Figure 4.15: Anion–dipole interaction.

Ion–dipole interactions are the strongest type of secondary interactions that will be considered in this course. They are the strongest as they exist between an ion (possessing a full negative or positive charge) and a neutral molecule (with a partial positive or negative charge). As solubility is dependent of the strength of the secondary interactions between solvents and solutes, molecules that are held together via an ion–dipole interaction are highly soluble in each other.

Science as a human endeavour

Ion–dipole interactions in the blood

The haemoglobin molecule contains four Fe^{2+} ions. Haemoglobin is the molecule responsible for the transportation of oxygen through the blood from the lungs to the tissues and cells that require it. In the early stages of the binding of an oxygen molecule to an Fe^{2+} ion in haemoglobin, it is the ion–dipole attraction that brings the O_2 towards the Fe^{2+}.

Questions

12. Write balanced ionic equations to represent the dissociation of the following salts. If the salt is insoluble in water, state "no dissociation".

Use Table 4.03 to help you determine the answers.

(a) silver chloride ($AgCl$)

....

(b) sodium acetate ($NaCH_3COO$)

....

(c) ammonium sulfate ($(NH_4)_2SO_4$)

....

(d) calcium carbonate ($CaCO_3$)

....

(e) potassium carbonate (K_2CO_3)

....

(f) sodium hydroxide ($NaOH$)

....

(g) iron(II) sulfate ($FeSO_4$)

....

(h) lead(II) phosphate ($PbSO_4$)

....

(i) lead(II) chromate ($PbCrO_4$)

....

(j) iron(III) chloride ($FeCl_3$)

....

(k) calcium nitrate ($Ca(NO_3)_2$)

....

(22 marks) **KA4**

13. Many fertilisers contain ammonium nitrate (NH_4NO_3).

(a) Write the balanced ionic equation to represent the dissociation of ammonium nitrate when it comes into contact with water.

.... *(2 marks)* **KA4**

(b) With reference to your equation above explain why ammonium nitrate is beneficial when used as a fertilizer.

....

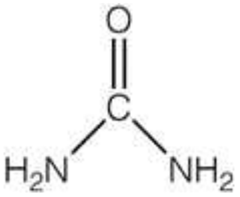
....

.... *(2 marks)* **KA1**

(c) Describe the likely intermolecular attraction that will arise when ammonium chloride dissociates in water.

(2 marks) **KA1**

(d) Some fertilisers also contain urea, as an inexpensive way of introducing nitrogen to the soil. The complete structural formula of urea is provided below.



Predict and explain whether urea is likely to be soluble or insoluble in water.

(2 marks) **KA1**

Precipitation reactions

Science understanding

Some ionic substances are not very soluble in water; such substances form as precipitates when solutions containing the relevant ions are mixed. *© Copyright SACE 2022*

Two water-soluble solutions can react to form an insoluble solid product. If the resulting product is insoluble it is called a **precipitate** (a solid which forms out of solution), resulting in a class of reactions known as **precipitation reactions**. They are useful in chemistry to assist in the identification of the presence of various ions in solutions.

Example

Lead(II) ions are toxic and may be present in polluted waterways. To test for the presence of lead(II) ions in contaminated water, sodium chloride solution can be added. The chloride ions combine with the lead(II) ions to form lead(II) chloride. This precipitates from the solution as a white solid. This is a positive test for lead(II) ions.

Precipitation reactions even occur in the human body between antibodies and antigens; however, the environment in which this occurs is still being studied.

Science as a human endeavour

Kidney stones

Kidney stones are solid masses made of crystals. They usually originate in the kidney but can be found anywhere along the urinary tract. They are very painful and can be medically serious.

Oxalate ions ($C_2O_4^{2-}$) are obtained from dietary sources and the metabolism of food within cells. These anions move to the kidneys where they combine with calcium ions (Ca^{2+}) to form a precipitate called calcium oxalate (CaC_2O_4). Calcium oxalate is a solid mass with an irregular crystal structure. The sharp surface of the crystal causes pain and inflammation within a person's urinary system. It is often suggested that a good step towards preventing kidney stones is to drink plenty of water. In essence, the solubility of the precipitate increases with the volume it is in – thus you can avoid the formation of these painful stones!

Precipitation, like crystallisation, is the reverse of dissociation. If a solid comes out of a solution slowly, a regular solid lattice has time to form and crystals are produced (crystallisation). If the solid is formed quickly, there is no time for a large regular lattice to form; instead, many small particles form in the liquid (precipitation). A common example is the combination of two clear solutions: silver nitrate and sodium chloride. By examining Table 4.03, it can be seen that both solutions are water-soluble; however, when placed in the same reaction, they undergo the following reaction.

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

The silver chloride is an insoluble salt, and therefore is classed as the precipitate. The prediction of such precipitation requires understanding and application of the **solubility rules**. Generally precipitation reactions are examples of single or double-replacement reactions. A double-replacement reaction occurs when two ionic reactants dissociate and bond with the respective anion or cation from the other reactant. The ions replace each other based on their charges as either a cation or an anion, as shown in Figure 4.16:

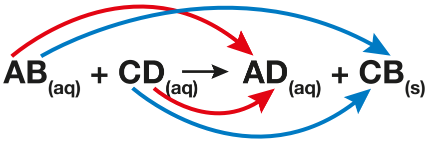


Figure 4.16: The movement of ions in a double-replacement reaction

The above balanced precipitation equation can be expanded to show all dissociated ions.

$$\mathbf{A^{+}_{(aq)} + B^{-}_{(aq)} + C^{+}_{(aq)} + D^{-}_{(aq)} \rightarrow A^{+}_{(aq)} + D^{-}_{(aq)} + CB_{(s)}}$$

This is referred to as a **complete ionic equation**. As the precipitate does not dissociate in water, the solid should not be separated. In the example above, the A^+ and D^- ions are shown on both sides of the equations. The A^+ and D^- ions do not undergo any chemical change during the reaction and are termed **spectator ions**. These species can be cancelled out in the same manner that elements are cancelled out from either side of a mathematics equation. The overall effect of this is called the **net ionic equation**.

$$\mathbf{C^{+}_{(aq)} + B^{-}_{(aq)} \rightarrow CB_{(s)}}$$

The net ionic equation is representative of the precipitation that has occurred, and is balanced in terms of the atoms of elements and the overall electric charge. If all products are aqueous, a net ionic equation cannot be written, as all ions would effectively be cancelled out.

The general steps for determining net ionic equations is shown below.

1. Write the balanced chemical equation, including the state of each substance.
2. Use the solubility table to rewrite the equation as the complete ionic equation, showing which compounds are dissociated into ions.
3. Cancel out the spectator ions that appear on both sides of the equation. The result is the net ionic equation.

Science as a human endeavour:

Soap scum

"Hard water" is a term often discussed in society when comparing the water quality in various locations around Australia. Chemically, hard water is water which contains higher than average concentrations of calcium and magnesium ions. An example of this is when calcium bicarbonate dissolves in water, the calcium and bicarbonate ions dissociate and become free in solution. As explained in Topic 4, soap and detergent ions have a negatively charged ionic carboxylate or sulfonate head. When the dissociated calcium ions come into contact with the carboxylate anion of soap ions, they are attracted to one another and bond.

$$Ca^{2+}_{(aq)} + 2RCOO^{-}_{(aq)} \rightarrow Ca(RCOO)_{2(s)}$$

The result is a salt that is no longer water-soluble, and it is this precipitate that is commonly called **soap scum**. Scum leaves unsightly deposits on clothes, in washing machines and in hair. In addition to this, the reaction renders the soap ions unable to form grease micelles to assist in the desired cleaning action. In comparison, calcium and magnesium sulfonates are soluble in water and thereby detergents do not form scum.

Questions

14. The “golden rain” reaction involves the addition of lead nitrate to potassium iodide. These clear and colourless solutions react and a golden coloured precipitate is formed. If the resultant solution containing the precipitate is heated and allowed to cool slowly, crystals of lead iodide form, causing the glistening “’golden rain” effect.

This reaction is summarised by the following equation:

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow 2KNO_{3(aq)} + PbI_{2(s)}$$

(a) For this reaction:

(i) Name the other product formed.

.. ..*(1 mark)* **KA1**

(ii) Complete the net ionic equation.

.. ..*(2 marks)* **KA4**

(iii) Identify the spectator ions.

.. ..*(1 mark)* **KA1**

(iv) Classify it as a precipitation or dissociation reaction.

.. ..*(1 mark)* **KA1**

(b) Explain why spectator ions are not included in a net ionic equation.

.. ..

.. ..*(1 mark)* **KA1**

(c) Describe the movement of ions in a double replacement reaction, using the reaction above to support your explanation.

.. ..

.. ..

.. ..*(2 marks)* **KA1**

15. Nickel(II) hydroxide, $Ni(OH)_2$ is used in nickel-cadmium batteries, as well as a catalyst in chemical reactions. It is the precipitate formed from the reaction between solutions of nickel(II) sulfate and sodium hydroxide.

(a) Identify the other product formed in the reaction.

.. ..*(1 mark)* **KA1**

(b) Write the complete ionic equation for the formation of nickel(II) hydroxide.

.. ..*(2 marks)* **KA4**

(c) Circle the spectator ions in the equation from (b)

(1 mark) **KA4**

(d) Complete the net ionic equation.

.. ..*(2 marks)* **KA4**

(e) Explain why this reaction is termed a double replacement reaction.

.. ..

.. ..*(1 mark)* **KA1**

16. For the double-replacement reactions shown below, complete the following:
    - (i) Deduce the products of each reaction
    - (ii) Write the complete ionic equation
    - (iii) Complete the net ionic equation

(a) $Fe(NO_3)_{3(aq)} + NaOH_{(aq)}$

(5 marks) **KA4**

(b) $Al_2(SO_4)_{3(aq)} + BaCl_{2(aq)}$

(5 marks) **KA4**

(c) $HI_{(aq)} + Zn(NO_3)_{2(aq)}$

(5 marks) **KA4**

4.3: Quantities in reactions

Chemistry can be broadly divided into two branches: qualitative and quantitative chemistry. Qualitative chemistry involves the description of a chemical reaction in terms of what change it undergoes which can be seen. For example, when magnesium burns in air, it produces a brilliant white flame, which results in the formation of magnesium oxide. Quantitatively this could be displayed as an equation:

$$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$$

Through stoichiometric techniques it could be determined that 10 g of magnesium burns to form 16.67 g of magnesium oxide. From this it can be deduced that quantitative analysis involves describing a chemical reaction in terms of the amounts (quantities) of substances in appropriate units. **Stoichiometry** is the term used to describe how, in chemistry, the relationships between reactants and/or products in a chemical reaction can be used to determine desired quantitative data. The term has Greek origins: *stokhein* translates to 'element', and *metron* means 'to measure', so together the term means 'the measure of elements'.

Quantitative chemistry is a very important branch of chemistry because it enables chemists to calculate known quantities of materials. Stoichiometry can determine how much product can be made from a known starting material, or how much of a given component is present in a sample.

Before a chemical company chooses to manufacture a product from a chemical process, stoichiometric calculations are performed to determine the amount of product that can be formed from known quantities of chemical reactants. Stoichiometry is fundamental to the decision-making processes of all chemical companies. If the predicted yield is too low, then the company can choose a different route (mechanism) for producing the product.

Quantitative analysis is any method used for determining the amount of a chemical in a sample. The amount is always expressed as a number with appropriate units. The mole concept (Topic 1.3) is central to quantitative chemistry. Chemists use the mole concept when determining the concentration of various solutions, and when examining the stoichiometry of precipitation reactions.

Chemical equations

Science understanding

Chemical equations can be written to describe a chemical change. *© Copyright SACE 2022*

A chemical equation is a symbolic expression of a chemical reaction. For example, when methane gas undergoes complete combustion in oxygen, carbon dioxide and water is formed. The unbalanced reaction equation for this reaction is given below.

$$CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$$

Chemical equations are used as the most efficient manner of displaying reactants and products and their ratios. When writing chemical equations, the following conventions are applied:

1. The correct formulae of the reactants are written on the left-hand side of the arrow.
2. The correct formulae of the products are written on the right-hand side of the arrow.
3. An arrow, $\rightarrow$, is used instead of the equals sign, =.
4. Like equations in mathematics, there must be the same number of each type of particle on the left-hand side as on the right (i.e. the number of atoms of each element on the left- and right-hand sides of the equation must be equal).
5. The sum of the electrical charges on the left- and the right-hand sides of the equation must be equal.
6. The coefficients written in front of the formulae indicate the number of particles (that is, atoms, molecules or formula units) of that substance in the reaction. Balancing equations is achieved by changing the coefficients in front of the formulae in the equation.

   Example - If $Al_2(SO_4)_3$ represents 2 aluminium atoms, 3 sulfur atoms and 12 oxygen atoms then

   $2Al_2(SO_4)_3$ represents 4 aluminium atoms, 6 sulfur atoms and 24 oxygen atoms
7. Polyatomic ions will usually remain intact from reactant to product, and therefore can be considered as one entity.

8. The physical states of the reactants and products under particular reaction conditions may be denoted by writing the following subscripts to the right of a chemical species:
    - solid – symbol is (s)
    - liquid – symbol is (l)
    - gas – symbol is (g)
    - dissolved in water – symbol is (aq).

Importantly, a chemical equation obeys the law of conservation of matter which states:

Matter is neither created or destroyed in a chemical reaction.

The total mass of the products of a chemical reaction must be equal to the total mass of the reactants. This concept was first identified by the Russian scientist Mikhail Lomonosov in 1756 and was confirmed by the French scientist Antoine Lavoisier in the decades that followed. An equation which conforms to the law of conservation of mass is said to be balanced.

When writing chemical equations, the following steps can be followed:

1. Write a word equation that includes all reactants and products.
2. Under each of the reactants and products in the word equation, write the correct formula of the species concerned.
    - This is of great importance, because if the formula of a species is incorrect, then the resultant equation will also be incorrect.
3. Alter the coefficients in front of each formula to balance the number of atoms of each element on both sides of the equation.

    You must not alter the:
    - subscripts that indicate how many of each atom is present in a substance (i.e. Na_2CO_3)
    - superscripts that indicate the charge of an ion (i.e. Al^{3+}).
4. Check that the number of atoms of each type of element is the same on both sides of the equation.
5. Write the physical states for each species.

The most effective way to understand and improve your ability to write chemical equations is to apply the above rules in example problems. Chemical equations are an important part of the chemical nomenclature that chemists apply, and it is only through practice that this skill develops. Some helpful examples are shown below:

Examples

Example 1

$$H_2 + Cl_2 \rightarrow HCl$$

In this example, both hydrogen and chlorine are examples of diatomic molecules, and therefore equal numbers of both must be present on the right-hand side for the equation to be balanced. In this instance, placing a coefficient of 2 for the hydrochloric acid balances the equation for completion:

$$H_2 + Cl_2 \rightarrow 2HCl$$

This happens fairly often at the end of a balancing sequence, when the placement of one coefficient balances two different elements at the same time.

Example 2

$$O_2 \rightarrow O_3$$

To balance this equation, the lowest common multiple (LCM) of 2 and 3 must be found, as it will indicate how many of each atom will be needed. In this case, the LCM is 6.

Once the LCM has been identified, the value of the LCM is divided by the value of the subscript for a given chemical species to determine the coefficient for that species.

i.e. In this instance the LCM is 6, and the subscript for O_2 is 2.

Therefore, to determine the coefficient, the LCM is divided by the subscript ($\frac{6}{2} = 3$), resulting in a coefficient of 3 for O_2.

Therefore, the correctly balanced equation is:

$$3O_2 \rightarrow 2O_3$$

Example 3

$$Na + H_2O \rightarrow NaOH + H_2$$

In the skeleton equation above, the Na and the O are already balanced, therefore attention is first given to the hydrogen molecule. H_2 is a diatomic molecule, and therefore an even number of hydrogen atoms must be present on the right-hand side. To achieve this, the following is done:

$$Na + H_2O \rightarrow 2NaOH + H_2$$

The left-hand side is then balanced by placing a 2 as the new coefficient for water:

$$Na + 2H_2O \rightarrow 2NaOH + H_2$$

Unfortunately, the result of this is that the sodium atoms are now unbalanced. To rectify this, a coefficient of 2 is placed in front of the sodium on the left-hand side. This completes the balanced equation:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Questions

17. Where necessary, balance the following chemical equations:

(a) $Al_{(s)} + O_{2(g)} \rightarrow Al_2O_{3(s)}$

(b) $CaI_{2(aq)} + Hg(NO_3)_{2(aq)} \rightarrow Ca(NO_3)_{2(aq)} + HgI_{2(s)}$

(c) $Mg_{(s)} + CH_3COOH_{(aq)} \rightarrow Mg(CH_3COO)_{2(aq)} + H_{2(g)}$

(d) $CoSO_{3(s)} \rightarrow CoO_{(s)} + SO_{2(g)}$

(e) $Fe_2O_{3(s)} + C_{(s)} \rightarrow Fe_{(s)} + CO_{(g)}$

(f) $Cu_{(s)} + S_{(s)} \rightarrow Cu_2S_{(s)}$

(g) $AgNO_{3(aq)} + K_2CrO_{4(aq)} \rightarrow Ag_2CrO_{4(s)} + KNO_{3(aq)}$

(h) $Sr(OH)_{2(aq)} + H_2SO_{4(aq)} \rightarrow SrSO_{4(s)} + H_2O_{(l)}$

(i) $Pb(NO_3)_{2(aq)} \rightarrow PbO + NO_{2(g)} + O_{2(g)}$

(j) $NH_{3(g)} + H_2SO_{4(aq)} \rightarrow (NH_4)_2SO_{4(aq)}$

(10 marks) **KA4**

18. Write and balance the following chemical equations:

(a) silver nitrate + magnesium chloride → magnesium nitrate + silver chloride

(b) zinc + cobalt(III) chloride → zinc chloride + cobalt

(c) iron(III) oxide + hydrochloric acid → iron(III) chloride + water

(d) sodium hydroxide + aluminum sulfate → sodium sulfate + aluminum hydroxide

(e) lead(II) nitrate + hydrochloric acid → lead(II) chloride + nitric acid

(f) copper sulfate + cadmium nitrate → cadmium sulfate + copper nitrate

(12 marks) **KA4**

Concentration of a solution

Science understanding

The concentration of a solution can be described in terms of mass concentration (mass of solute per unit volume, ρ) or as molar concentration (moles of solute per unit volume, c).

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Concentration is the general term used in referring to the quantity of solute in a standard quantity of solution. A solution is considered **dilute** when the concentration of solute per unit volume is low, and considered **concentrated** when the solute to volume ratio is high. The more concentrated is solution is, the more particles it possesses per given unit of solution.

To obtain a precise quantity of solute, quantitative methods of expressing concentration are required. This is essential in chemical analysis, chemical manufacturing, pharmaceutical sales, and many other facets of modern society.

Two methods of expressing solution concentration are in terms of **mass concentration** (mass of solute per unit volume, ρ) or **molar concentration** (moles of solute per unit volume, c).

Molar concentration

The concentration of a solution in moles per litre is sometimes referred to as the molar concentration, or **molarity** of a solution, and given the symbol *c*. It is the number of moles of a solute dissolved in 1 litre of solution. It is quantitatively determined via the following equation:

$$\text{Molarity (mol L}^{-1}) = \frac{\text{number of moles of solute}}{\text{number of litres of solution}} \text{ and simplified to } c = \frac{n}{V}$$

Examples

Example 1

A 500 mL solution is prepared by adding 0.2 moles of sodium hydroxide to a 500 mL volume of distilled water. Calculate its concentration.

$V = 500 \text{ mL } (0.5 \text{ L})$

$$c = \frac{n}{V} = \frac{0.2}{0.5} = 0.4 \text{ mol L}^{-1}$$

Example 2

A saline solution contains 0.9 g of NaCl per 100 mL of solution. Calculate its concentration.

$$n_{(NaCl)} = \frac{m}{M} = \frac{0.9}{58.44} = 0.0154 \text{ mol L}^{-1}$$

$V = 100 \text{ mL } (0.1\text{L})$

$$c = \frac{n}{V} = \frac{0.0154}{0.1} = 0.15 \text{ mol L}^{-1}$$

Such calculations are examples of **a mass-volume calculation**. This means either the mass of a substance is given and the volume of another is calculated, or the volume of a substance is provided and the mass of another compound is calculated.

Mass concentration

The mass of solute per unit volume (or volumetric mass density) is chemically determined via the following formula:

$$\rho = \frac{m}{V}$$

ρ refers to the calculated density in kg/m^3 (or equivalently for chemistry, g/cm^3), m is the mass of solute, and V is the volume of solution in which the solute is placed.

More commonly in chemistry, mass concentration is calculated in g L^{-1} (which is an equivalent unit to kg/m^3, g/cm^3 and mg/mL), and shown via the following formula:

$$\text{Concentration (g L}^{-1}) = \frac{\text{mass of solute (in grams)}}{\text{volume of solution (in litres)}} \text{ and simplified to } c = \frac{m}{V}$$

Example

Calculate the mass concentration of a solution containing 30 g of sodium bicarbonate dissolved into a 1 L volume of distilled water?

V = 1L

$$c = \frac{m}{V} = \frac{30}{1} = 30 \text{ g L}^{-1}$$

Representing mass concentration as g L^{-1} is not always the most relevant unit for communicating the mass concentration of a solution. With this in mind, other conventions such as percent weight per volume ($\% \frac{w}{V}$), parts per million (ppm) and parts per billion (ppb) can be investigated.

Percent weight per volume

The percent weight per volume represents the mass of solute in grams, dissolved in 100 millilitres (mL) of solution. This is often the unit of choice or consumer products such as household cleaning products and medicinal products.

This is represented by the symbol $\% \frac{w}{V}$ and the unit g.100 mL^{-1}.

The expression for the calculation of percent weight per volume is:

$$\text{percent weight per volume} = \frac{\text{mass of solute (in grams)}}{\text{volume of solution (in millilitres)}} \times 100$$

$$\% \frac{w}{V} = \frac{m_{\text{solute}}}{V_{\text{solution}}} \times 100$$

Parts per million

The number of parts per million depicts the mass of solute, measured in milligrams (mg), dissolved in 1.00 kilogram (kg) of solution. The term parts per million is used as there are one million milligrams in one kilogram, therefore one milligram per kilogram is one part in a million parts. This is the unit of choice for solutes with very low concentrations, such as the salts present in mineral waters or heavy metal pollutants in rivers or lakes.

As the density of water is 1.00 g mL^{-1}, it follows that the density of a very dilute aqueous solution is approximated as 1.00 g mL^{-1}. Consequently, a solution containing 1.00 mg of solute per 1.00 litre of solution, or 1.00 mg L^{-1}, contains the equivalent of 1.00 mg of solute per 1.00 kg of solution. This is equivalent to a concentration of 1.00 ppm.

For very dilute aqueous solutions:

1.00 g of solution is equivalent to 1.00 mL of solution

1.00 kg of solution is equivalent to 1.00 L of solution

The symbol for parts per million is ppm with the unit of mg L^{-1}. It is expressed as:

$$\text{parts per million} = \frac{\text{mass of solute (in milligrams)}}{\text{volume of solution (in litres)}}$$

$$\text{ppm} = \frac{m_{\text{solute}} \text{ (mg)}}{V_{\text{solution}} \text{ (L)}}$$

Parts per billion

The number of parts per billion represents the mass of solute, measured in micrograms (μg), dissolved in 1.00 L of solution. This is the unit of choice for solutions containing trace quantities of solutes, such as those found in large bodies of water.

The symbol for microgram is μg (1μg = 1.00×10^{-6} g)

Parts per billion is presented by the symbol ppb and measured via the unit μg L^{-1}. It is expressed by:

$$\text{parts per billion} = \frac{\text{mass of solute (in micrograms)}}{\text{volume of solution (in litres)}}$$

$$\text{ppb} = \frac{m_{\text{solute}} \text{ (μg)}}{V_{\text{solution}} \text{ (L)}}$$

Conversions of units

At times it is useful to change, or 'convert', concentration units from one to another. In terms of mass and molar concentration, this conversion is relatively simple and shown in Figure 4.17.

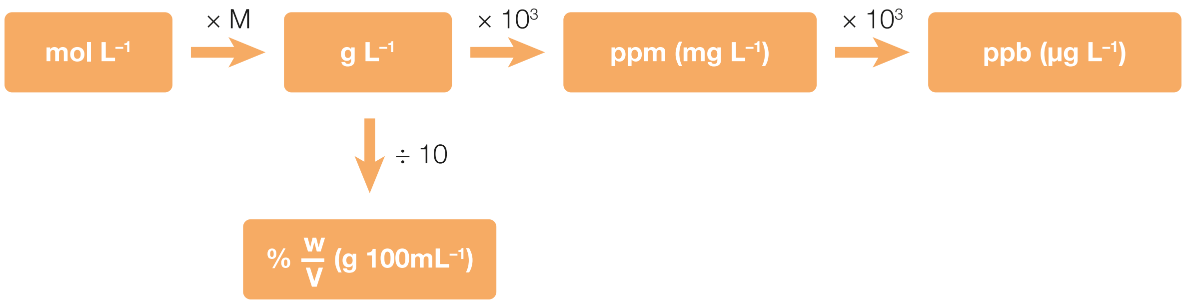


Figure 4.17: The relationships between different solution concentration units

Questions

19. Silver nitrate, $AgNO_3$ is useful in the cauterisation of wounds and ulcerations, as well as the removal of warts. Aqueous solutions of silver nitrate are also useful in analytical chemistry for the determination of halides and cyanides.

(a) Calculate the number of moles present in a 0.9 g sample of $AgNO_3$.

(2 marks) **KA4**

(b) Determine the concentration of solution if this sample of $AgNO_3$ is dissolved into 250.00 mL of distilled water.

(2 marks) **KA4**

(c) Convert this concentration to g L^{-1}.

(2 marks) **KA4**

(d) The solution produced is said to be dilute. Explain what is meant by this term.

(1 mark) **KA1**

20. A student conducts an investigation into the salinity levels of the ocean which arises from the presence of sodium and chloride ions.

As one part of the study, a solution is prepared to mimic the conditions of the seawater. The student dissolves 16g of sodium chloride in 500mL of distilled water.

(a) Determine the concentration of the student's solution in g L^{-1}.

(1 mark) **KA4**

(b) Using a salinity probe and data logger the student then tested the salinity level of a sample of seawater. It was determined to have a concentration equal to 35,000 ppm.

Express this concentration in mol L^{-1}

(3 marks) **KA4**

(c) State and explain if the student successfully created a solution that mimicked the salinity levels of seawater.

(2 marks) **KA1 and KA4**

21. (a) Calculate the volume of 2.0 mol L^{-1} hydrochloric acid needed to completely dissolve 5.4 g of aluminium.

$2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

(b) Calculate the mass of hydrogen that would be evolved.

(11 marks) **KA4**

22. The concentration of glucose ($C_6H_{12}O_6$) in the blood of a person was determined by analysis to be 4.5×10^{-3} mol L^{-1} . Convert this concentration to:

    (a) Mass concentration (g L^{-1})

    (b) Mass concentration (ppm)

(5 marks) **KA4**

23. The label of spring water states that the water contains magnesium at a concentration of 4 mg L^{-1}. Convert this concentration to:

    (a) % $\frac{w}{V}$ concentration

    (b) ppb concentration

(6 marks) **KA4**

Dilution

Laboratory chemicals, industrial solvents and food products are often provided to the consumer in a concentrated form. One of the many important laboratory techniques is the dilution of these concentrated solutions into a concentration that would be suitable for consumption (in the case of pharmaceuticals or alcoholic beverages).

In the process of dilution, it is important to recognise that the number of moles of the original concentrated solution will remain the same as that in the diluted solution, however the solute is distributed through a greater volume of solution.

The initial moles of a solution can be calculated via:

$$n_{initial} = c_{initial} \times V_{initial}$$

And the moles of the final by:

$$n_{final} = c_{final} \times V_{final}$$

Therefore, if the moles of the initial equal that of the final, the equation can become:

$$c_{initial} \times V_{initial} = c_{final} \times V_{final}$$

The above equation can be rearranged to determine any of the above four variables, given that the other three are known.

Such calculations are of particular importance in the preparation of standard solutions. A standard solution is a solution of accurately known concentration. It is prepared from a primary standard (a compound which is stable, of high purity, highly soluble in water and of a high molar mass to allow for accurate weighing) that is weighed accurately and made up to a fixed volume.

Standard solutions are prepared via the following steps:

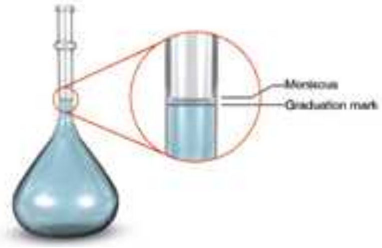
- Decide desired concentration and volume of standard.
- Calculate moles ($n = c \times V$) and then mass ($m = n \times M$) of primary standard needed.
- Use a mass balance to accurately weigh and record a mass of primary standard that approximates the pre-calculated mass.

- Using a funnel, transfer standard to a washed and rinsed volumetric flask, ensuring that all traces of the standard are thoroughly rinsed with distilled water from the beaker and funnel.

Swirl to ensure the solid solute has dissolved.

Fill the volumetric flask so that the bottom of the meniscus is on the calibration mark.

Stopper and invert the solution just once to mix the contents of the flask and create a homogenous solution. Inverting the flask multiple times dissolves air (trapped between the lid and meniscus) into the solution and gives an inaccurate reading of the true volume of solution.



Questions

24. Sodium hydroxide is a commonly prepared standard solution as it supplied to laboratories in highly concentrated form. Three students prepared 500 mL standard solutions from differing stock solutions for the purpose of their investigations.

   (a) For each student below, calculate how many millilitres of stock solution would be required to make standard solutions of concentration 0.2 mol L^{-1}.

      (i) Student 1 Stock Solution: 0.25 mol L^{-1} sodium hydroxide.

      (2 marks) **KA4**

      (ii) Student 2 Stock Solution: 1.0 mol L^{-1} sodium hydroxide.

      (2 marks) **KA4**

      (iii) Student 2 Stock Solution: 2.5 mol L^{-1} sodium hydroxide.

      (2 marks) **KA4**

   (b) Describe how Student 1 would prepare their standard solution.

   (2 marks) **KA1**

   (c) After the completion of their investigations the students were posed with further calculations to solve. These are shown below, complete each calculation that is posed.

      (i) Determine the volume of concentrated 18 mol L^{-1} stock sulfuric acid needed to make 2.00 L of 0.200 mol L^{-1} solution by dilution of the original concentrated solution.

      (2 marks) **KA4**

      (ii) Calculate the volume of 15.8 mol L^{-1} nitric acid solution that is needed to make 4.25 L of 0.250 mol L^{-1} nitric acid solution.

      (2 marks) **KA4**

25. A 250.0 mL sodium carbonate standard solution was prepared from distilled water and 21.2 g of the solid.

    Describe the correct steps and equipment (including the final concentration) that would be required to prepare the above solution.

    (8 marks) **KA1**

Stoichiometric calculations for precipitation reactions

Science understanding

Chemicals react in definite proportions.

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As introduced earlier in this workbook, stoichiometry is a quantitative process. Given an initial mass or volume and concentration of reactant or product, the molar relationships between reactants and products in a chemical reaction are used to calculate a specific mass or volume of another reactant or product.

When reactions are carried out in a laboratory, quantities are measured in grams, not moles. It is for this reason that most calculations require you to start and finish with mass. This is specifically termed **mass–mass stoichiometry**. Mass–mass stoichiometry can be combined with knowledge of precipitation reactions to quantify the amount of salt present in a solution. When this quantitative analysis is being undertaken, knowledge of the solubility rules is important and the use of a solubility table is necessary.

To solve mass–mass stoichiometry problems the following steps are followed:

1. Write a balanced chemical equation for the reaction.
2. Calculate the moles of the precipitate from its mass, using $n = \frac{m}{M}$
3. Use the mole ratio to find the number of moles of the salt in solution
4. Calculate the mass of the salt using $m = n \times M$

Example

A reaction between solutions of sodium sulfate and barium nitrate produces a precipitate of barium sulfate with a mass of 2.44 g. Calculate the mass of sodium sulfate required to produce the barium sulfate precipitate.

$$Na_2SO_{4(aq)} + Ba(NO_3)_{2(aq)} \rightarrow BaSO_{4(s)} + 2NaNO_{3(aq)}$$

$n = \frac{m}{M} = \frac{2.440}{233.4} = 0.0105$ mol.

Using mole ratios = $\frac{\text{unknown}}{\text{known}}$, $\frac{Na_2SO_4}{BaSO_4}$, $\frac{1}{1}$

Therefore the moles of Na_2SO_4 are equal to 0.0105 mol.

m = n × M , = 0.0105 × 142.04 = 1.48 g

Questions

26. A blue coloured solution is produced when 17.0g of copper sulfate is dissolved in water. When a strip of magnesium metal is placed in the solution, the blue colour of the disappears and the magnesium metal turns brown as the displaced copper is deposited on it. Along with solid copper, aqueous magnesium sulfate is also produced.

    (a) Write the balanced chemical equation for this reaction.

    .. ..*(2 marks)* **KA4**

    (b) Determine the number of moles present in the 17.0g sample of copper sulfate.

    (2 marks) **KA4**

    (c) Using mole ratios, determine the number of moles of copper present at the end of the reaction.

    (2 marks) **KA4**

    (d) Calculate the mass of copper that will be produced.

    (2 marks) **KA4**

27. A spontaneous reaction takes place when aluminium metal is placed in a solution of zinc chloride resulting in the formation of zinc metal.

    (a) Along with zinc metal, an aqueous salt is formed. Name the salt produced.

    .. ..*(1 mark)* **KA1**

    (b) Write the balanced chemical equation for the reaction described above.

    .. ..*(2 marks)* **KA4**

(c) The resultant solution was evaporated to regain the salt produced, it was weighed and determined to have a mass of 14.7g.

(i) Calculate the number of moles present in the 14.7g sample of salt.

(2 marks) **KA4**

(ii) Using mole ratios, determine the number of moles present in the zinc chloride solution at the start of the reaction.

(2 marks) **KA4**

(iii) Determine the mass of zinc chloride that was dissolved in water before the aluminium was added.

(3 marks) **KA4**

28. A reaction between solutions of potassium iodide and lead (II) nitrate is performed, resulting in 1.67 g of lead (II) iodide precipitate. Calculate the mass of potassium iodide required to produce this.

(9 marks) **KA4**

29. Solutions of sodium chloride and silver nitrate react to produce 0.115g of silver chloride. The reaction is represented by the following equation:

$$NaCl_{(aq)} + AgNO_3 \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

Calculate the mass of sodium chloride that must have been present to produce this precipitate.

(5 marks) **KA1**

4.4: Energy in reactions

Chemical reactions involve the transformation of reactants to products. This can be marked by a change in colour, temperature or other chemical property. On an atomic level, this transformation is due to the rearrangement of atoms that can only take place if there is sufficient energy present for the reaction to proceed. In order to effectively initiate a reaction, collisions between reactants must be sufficiently **energetic**. In other words, energy is a requirement to initiate a chemical process.

This is linked intrinsically to the collision theory of chemical reactions. This theory is based on the premise that, for a reaction to occur:

- the particles (molecules, atoms or ions) of which the reactants are composed must collide
- the colliding particles must have sufficient energy to form an activated complex, an unstable intermediate from which the final product is formed.

The **activation energy** (E_a) of a reaction is the minimum energy required for reactant particles to form the activated complex.

Chemical reactions are accompanied by the gain or loss of energy

Science understanding

All chemical reactions involve the formation of a new substance and are accompanied by the gain (endothermic reactions) or loss of energy (exothermic reactions).

The energy released in endothermic and exothermic reactions can be explained in terms of the Law of Conservation of Energy and the breaking and forming of bonds.

In accordance to the law of conservation of energy, the total energy of a system must remain unchanged in a chemical reaction. The energy change in a chemical reaction is due to the difference in the amounts of stored energy between the products and reactants. Often a reaction will absorb or release energy in forms such as sound, electrical or electromagnetic radiation (heat, light or both). This stored chemical energy, or heat content of the system, is referred to as **enthalpy**. In reactions where the system *releases* heat to the surroundings, the reactants have more stored energy (greater total enthalpy) than the products. This class of reaction is termed **exothermic**. This type of reaction is one where there is a net energy release to the surroundings. The surroundings include:

- water (if the reaction occurs in water)
- air (i.e. combustion/reaction with oxygen)
- the reaction vessel.

Science as a human endeavour:

Photosynthesis and Aerobic Respiration

Photosynthesis

Vital for life on Earth, the process of photosynthesis provides an example of an endothermic reaction. The process by which plants, algae and phytoplankton convert carbon dioxide and water into glucose and oxygen, occurs via a series of reactions, however is it summarised by this following overall thermochemical equation.

$$\Delta H = +2802 \text{ kJ mol}^{-1}$$

Unlike most other endothermic reactions, the energy absorbed for the reaction to proceed does not come from heat. The plants absorb light, UV energy radiated from the sun, and use this to allow the reaction to proceed.

Aerobic Respiration

The macromolecules that we consume in our food are broken down by the body, providing usable energy. Energy is essential for cellular reactions and processes that continually take place. These reactions allow us to move, grow, repair, our cells to divide, and to maintain homeostasis.

Carbohydrates, a class of macromolecule, are firstly broken down into their monomer units, the most common of which is glucose. It is then the subsequent oxidation of glucose, that releases energy that is used by the body to do work. This process is termed aerobic respiration. Like photosynthesis, aerobic respiration involves a series of complex reactions, however can be summarised by an overall thermochemical reaction. The equation of aerobic respiration is shown below.

$$\Delta H = -2805 \text{ kJ mol}^{-1}$$

As the intramolecular bonds of glucose are broken via this process, energy is released and the process is exothermic.

In certain reactions, the system may *absorb* energy from the surroundings, causing the temperature of the surrounding to decrease. In these instances, the reactants have less stored energy (less total enthalpy) than the products. The reaction is classed as **endothermic** as it involves a net absorption of energy from the surroundings.

It is possible to measure the changes in internal energy that occur during a chemical reaction. As most reactions take place in vessels that are open to the atmosphere, they take place at constant pressure (even though the volume may vary). The change in heat energy content measured at constant pressure is **enthalpy change** and is represented by the symbol ΔH.

Enthalpy (H) is a measure of the internal energy of a system (U) at constant volume and pressure. If pressure and volume are constant, the change in enthalpy (ΔH) is proportional to changes to the internal energy of the system (U). If energy enters a system from the surroundings, the internal energy of the system (U) increases. The change in enthalpy is positive (endothermic) as there has been an increase in internal energy. When there is a net energy release from the system to the surroundings, the internal energy of the system decreases. The internal energy (U) and enthalpy (H) decrease. The enthalpy change is therefore negative (exothermic). This results in the following:

The value of ΔH is given a negative sign for an exothermic reaction.

The value of ΔH is given a positive sign for an endothermic reaction

This indicates if the products have gained or lost energy to the surroundings. Another way to describe ΔH is the difference between the total energy of the products and the total energy of the reactants.

Thus, if the products have less energy than the reactants, ΔH will be negative (exothermic) and energy will be released from the reaction.

The opposite is also true. Therefore ΔH can be expressed as:

$$\Delta H = H_{products} - H_{reactants}$$

The enthalpy change for a reaction is measured in joules (J) or kilojoules (kJ) per mole or gram of specified reactant or product. The unit kJ mol^{-1} may be converted to kJ g^{-1} by dividing by the molar mass.

The molar enthalpy change for a reaction is the quantity of heat energy released or absorbed when 1 mole of a specified substance reacts in a chemical reaction under constant pressure. It is quoted in kJ mol^{-1} of the specified chemical.

Reaction equations for which enthalpy change is specified are written as **thermochemical equations**. This class of equation is unique as such equations provide information about the relationships between the quantities of the reactants and the enthalpy change (ΔH) for the reaction.

They include:

- The mole ratios of reactants and products.

  For clarity, the equation is usually balanced by using only one mole of the subject substance.
- The state of each reactant and product.
- The quantity of heat energy released or absorbed by the mole quantities of reactants or products indicated in the equation.

Example

$$MgCO_{3(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)} \quad \Delta H = -90.4 \text{ kJ mol}^{-1}$$

This equation conveys the following information:

When 1.00 mol (83.32 g) of solid magnesium carbonate reacts with 2.00 mol of aqueous hydrochloric acid, the following form:

1.00 mol of magnesium chloride in aqueous solution, 1.00 mol of carbon dioxide and 1.00 mol of liquid water.

90.40 × ***y*** kJ of heat energy is released to the surroundings.

*It follows from this equation that **y moles** of solid magnesium carbonate reacting in this way would release 90.40 × y kJ of heat energy to the surroundings, or that one gram would release* $\frac{90.4}{84.32}$ *kJ of heat energy.*

Questions

30. For each reaction described below:

    (i) Classify the reaction as exothermic or endothermic.

    (ii) Write the thermochemical equation for the reaction.

    (a) Methane (CH_4) gas is used as a fuel for heating hot water in many of our homes. In addition, it is the gas used to fuel the Bunsen burners in our laboratories. The ΔH(combustion) for methane is –890 kJ mol.

....

.... *(5 marks)* **KA1 and KA4**

    (b) During the synthesis of water, hydrogen and oxygen combine. This results in the release of 286 kJ of energy.

....

.... *(5 marks)* **KA1 and KA4**

    (c) Photosynthesis requires an energy input of +2802 kJ mol^{-1} to produce glucose and oxygen.

....

.... *(5 marks)* **KA1 and KA4**

Dissolving ionic substances in water (enthalpy of solution)

Science understanding

When ionic substances dissolve in water the dissociation of ions requires energy and the hydration of the ions releases energy. *© Copyright SACE 2022*

The **enthalpy of solution**, also referred to as the heat of solution or enthalpy of dissolution, is associated with the dissociation of an ionic solute in a solvent. It is expressed in kJ mol^{-1}, and must occur under constant pressure.

Such enthalpy is classed as exothermic or endothermic and is dependent on two energetic processes that occur when the ionic salt lattice is placed into the solvent (in most cases water) and dissociates completely into its resultant ions.

The first is termed **lattice energy**. This is the energy that is applied to 'pull' the ions away from each other in the lattice. Lattice energy is central to the dissociation of the salt.

Once the ions are successfully dispersed in a solution, energy is released into the solution as water molecules surrounding the ions. This process is referred to as the **hydration energy**. Lattice energy and hydration energy are depicted in Figure 4.18

Figure 4.18: Lattice and hydration energy of sodium chloride

Essentially, the enthalpy of solution for an ionic compound is dependent on which of the two energies (lattice or hydration) is greater. Using the above figure of sodium chloride as an example, the lattice energy is 779 kJ mol^{-1} and its hydration energy is 774 kJ mol^{-1}. As more energy is involved in the dissociation of the sodium chloride lattice, compared to the energy released when new secondary interactions are formed between ions and water, this reaction is endothermic. This process is represented by the following thermochemical equation and simplified expression:

$$NaCl_{(aq)} + aq \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)} \qquad \Delta H_{solution} = +5 \text{ kJ mol}^{-1}$$

$$\Delta H(\text{solution}) = \text{lattice energy} - \text{hydration energy}$$

When the lattice energy exceeds that of the hydration energy, the reaction is **endothermic**.

If it is the hydration energy that is greater than the lattice energy, the reaction is **exothermic**.

The process of dissociation (dissolving) is exothermic when more energy is released when water molecules "bond" to the solute, than than is used in the dissociation of the anions and cations. As more energy is released than is used, the molecules of the solution move faster, making the temperature of the water increase. Dissolving potassium hydroxide in water is exothermic. This is because more energy is released upon formation of solute–solvent bonds between the ions and water molecule than was required to break apart the hydrogen bonds in water and the ionic bonds in potassium hydroxide.

Lattice energies and hydration energies of ionic compounds are known values, and therefore can be sourced to determine the overall enthalpy of solution. Such values are shown in Table 4.04.

Salt	Lattice energy (kJ mol^{-1})	Hydration energy (kJ mol^{-1})
NaCl	779	774
KCl	701	685
LiF	1032	1005
NaOH	737	799
AgCl	916	851
$SrCl_2$	2110	2161
KI	632	617

Table 4.04: Standard lattice and hydration energies of common salts.

Science as a human endeavour:

Hot and cold packs

Hot and cold packs are used by athletes to minimise swelling of injuries such as muscle and joint sprains. Such packs are constructed of a large pouch containing a dry chemical compound and an inner pouch of water. The hot or cold pack is activated by breaking the seal on the pouch of water and shaking the pack vigorously. This action mixes the water with the chemical compound and initiates the exothermic or endothermic reaction.

Commercial instant cold packs typically use either ammonium nitrate or urea as the salt; hot packs often use magnesium sulfate or calcium chloride. These reactions happen in a similar manner. When the salt is dissolved in water, the ionic bonds are overcome. This process requires energy, which is obtained from the surroundings. The ions then form bonds with the water, a process that releases energy. If more energy is released than taken in, then the process is exothermic, making the solution feel warmer. If more energy is taken in than released, then the process is endothermic, making the solution feel cooler.

Questions

31. Cold packs are often used in the treatment of sporting injuries. Immediate application of a cold pack has been shown to reduce swelling after an acute injury such as a sprain.

Some cold packs contain ammonium chloride (NH_4Cl), separated from water in the cold pack by a thin membrane. When struck by our hands, the membrane breaks and the ammonium chloride dissolves in the water.

The reaction is represented by the following thermochemical equation:

$$NH_4Cl_{(s)} + aq \longrightarrow NH^{+}_{4(aq)} + Cl^{-}_{(aq)}\ \Delta H_{(solution)} = +17 \text{ kJ mol}$$

(a) Identify if this reaction is endothermic or exothermic.

.. ..*(1 mark)* **KA1**

(b) State if energy is released or absorbed by this reaction.

.. ..*(1 mark)* **KA1**

(c) The enthalpy change of this reaction is positive. State what this tells us in terms of the energy of reactants and products.

.. ..

.. ..*(1 mark)* **KA1**

(d) Energy is required to initiate the chemical process. Explain why.

(2 marks) **KA1**

(e) Explain how would you apply the law of conservation of energy to the situation with the cold pack.

(2 marks) **KA1**

(f) The lattice energy of NH_4Cl is equal to 705 kJ mol^{-1}. Calculate the compounds hydration energy.

(1 mark) **KA1**

(g) With reference to your answer above, explain the difference between lattice and hydration energy.

(3 marks) **KA1**

32. (a) Complete the table by calculating the enthalpy of solution of for each of the following compounds. Refer to Table 4.04 to assist with this.

Salt	Enthalpy of Solution (∆H)
NaCl	
KCl	
LiF	
NaOH	
AgCl	

(5 marks) **KA4**

(b) The enthalpy of solution for sodium iodide is –7.5 kJ mol^{-1}. Describe what this means in terms of its dissociation in water.

(3 marks) **KA1**

Enthalpy changes

Science understanding

Enthalpy changes for solution reactions can be determined experimentally.

The measurement of heat changes in a chemical reaction is referred to as **calorimetry**. The insulated apparatus used to measure these heat changes is called a calorimeter. For a calorimeter to be useful it must be well insulated so that any energy gain or loss to the surrounding environment is minimised. The temperature change of water is used to measure the energy released or absorbed by the reaction. A calorimeter can vary in complexity from a simple Styrofoam cup with a cardboard lid to a sophisticated bomb calorimeter that has thick steel walls maintaining a high pressure environment. A simple and bomb calorimeter are presented in Figure 4.19.

Figure 4.19: A simple calorimeter compared to a complex calorimeter.
Please note: the bomb calorimeter is used for applications where the fuel is combusted.

Upon viewing their designs it is evident that the construction of the simple calorimeter would introduce systematic errors. So, when conducting calorimetry, there are assumptions that are adhered to:

- All heat is transferred to the water inside the calorimeter (there is no heat loss between the contents and the atmosphere).
- The maximum or minimum temperature is reached before any heat has been lost or gained from the atmosphere.

The change in enthalpy of a system is proportional to the change in temperature of that system at constant pressure. Therefore, the change in enthalpy for a given reaction can be calculated by measuring changes to the temperature of the system at constant pressure. To facilitate the conversion of temperature from degrees Celsius to demonstrate the heat absorbed or released by a reaction (measured in J), the following equation is used:

$$Q \text{ (heat energy absorbed or released)} = m \text{ (water)} \times \Delta T \times c_p \text{ (water)}$$

Where:

- Q is the energy released or absorbed in the reaction
- m is mass of water in grams in the calorimeter
- ΔT is the change in temperature of the solution
- c_p is the specific heat capacity of the solution (for water this is 4.18 J g^{-1} $°C^{-1}$). The specific heat capacity is a constant that refers to the amount of energy in joules required to increase the temperature of 1 g of water by 1°C.

Further to this, the heat absorbed or released can be then used to determine molar enthalpy of the reaction. The expression for Q is used as a base for the development of the subsequent equation:

$$\Delta H \text{ (enthalpy change)} = \frac{m_{water} \times \Delta T_{water} \times c_{p\ (water)}}{1000 \times n_{substance}}$$

Note: dividing by 1000 converts the unit from joules to kilojoules.

And simplified to:

$$\Delta H \text{ (enthalpy change)} = \frac{Q}{1000 \times n(\text{substance})}$$

$$\Delta H \text{ (kJ mol}^{-1}\text{)} = \frac{Q \text{ (kJ)}}{n\text{(mol)}}$$

Enthalpy change is measured in the unit kJ mol^{-1}. Along with the assumptions mentioned above, there are also several approximations made in the mathematical analysis of enthalpy. Like the assumptions, these approximations are responsible for introducing sources of systematic errors. These are:

The specific heat capacity of any solution is the same as water (4.18 J g^{-1} $°C^{-1}$)

The density of any solution is the same as water (1.0 g mL^{-1})

If the enthalpy of reaction is already known and a specific quantity is given, then the following equation can be used to determine the heat of reaction. This value is measured in kJ.

$$Q = n_{substance} \text{ x } \Delta H_{substance}$$

Example

An experiment was conducted in which 5.19 g of sodium carbonate was dissolved in 75.0 g of distilled water.

$$Na_2CO_{3(s)} + aq \rightarrow 2Na^{+}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$$

A temperature increase of the system of 3.80°C was observed. Calculate the heat absorbed or released for this reaction.

$Q = m(water) \times \Delta T \times c_{p\ (water)}$

$-Q = 75.0 \times 3.80 \times 4.18$

$-Q = 1191.3$ J

$Q = -1200$ J (2s.f) (as heat is released Q is negative)

The resultant enthalpy can then be determined:

$$\Delta H = \frac{Q}{1000 \times n(substance)}$$

$$\Delta H = \frac{-1191.3}{\frac{5.19}{105.99} \times 1000}$$

$\Delta H = -24.3$ kJ mol^{-1} (3s.f)

In applications of all the above equations, it is important to consider significant figures. When considering significant figures, the following guidelines are applied:

1. **All non-zero digits (integers) are significant**
2. **Zeros between non-zero digits are significant**
3. **Zeros at the end of a number are significant**
4. **Zeros to the left of the first non-zero digit are not significant**
5. **Zeros at the end of a number without a decimal point are ambiguous.**

In chemistry, answers are generally provided to the number of significant figures of the least accurate piece of data used in the calculation.

Questions

33. (a) Calculate the number of joules:

(i) Needed to warm 25.5 grams of water from 14ºC to 22.5ºC.

(2 marks) **KA4**

(ii) Released when 75.0 grams of water are cooled from 100ºC to 27.5ºC.

(2 marks) **KA4**

(b) Explain what is meant by the term specific heat capacity.

(2 marks) **KA1**

34. A 7.0 g sample of caesium is sealed in a glass vial and lowered into 250.0 mL of water at 90.00°C. When the caesium had melted, the temperature of the water had dropped to 88.98°C. Determine the molar enthalpy for caesium.

(4 marks) **KA4**

35. If 8.0 g ammonium nitrate is dissolved in 1.0 L water, the water decreases in temperature from 21.0°C to 20.39°C. Determine the molar heat of solution of the ammonium nitrate.

(4 marks) **KA4**

36. The enthalpy of reaction between magnesium metal and hydrochloric acid was to be quantified in a laboratory.

A 0.1277 g piece of magnesium ribbon was added to 200.0 mL 0.500 mol L^{-1} HCl at 24.12 °C. The water temperature increased to 27.10 °C.

(a) Calculate the enthalpy change for this reaction.

(4 marks) **KA4**

(b) Describe what assumptions are made in calorimetry.

(2 marks) **KA1**

(c) List the sources of systematic errors in calorimetry.

(3 marks) **KA1**

(d) Students in a class conducted the same experiment but due to the inherent systematic errors all calculate differing values for the enthalpy of reaction. These are shown in the below table.

Indicate the number of significant figures shown in each student's value.

	Enthalpy of Reaction (-ΔH)	**Number of Significant Figures**
(i)	474	
(ii)	487.6	
(iii)	400	
(iv)	400.65	

(4 marks) **KA4**

Summary Test 4: Mixtures and solutions

1. Two unlabelled samples are provided with the experimental aim to determine which substance is cyclohexane and which is ethanol.

   (a) Using the principle of miscibility (with water as the solvent) describe an experimental technique that could be used to distinguish the two substances.

   (2 marks) **KA4**

   (b) Use a diagram to explain the relationship between a solvent and solute.

   (3 marks) **KA4**

   (c) Explain what secondary interactions exist between the molecules of ethanol and water when they mix.

   (2 marks) **KA1**

   (d) The mixture formed between ethanol and water would be described as a homogenous solution. Describe what is meant by this term.

   (2 marks) **KA1**

   (e) Explain how miscibility is dependent on molecule polarity.

   (2 marks) **KA1**

   (f) Draw a labelled diagram of ethanol indicating regions of polarity.

   (2 marks) **KA4**

(g) Like ethanol, hexan-1-ol also possesses a polar hydroxyl group, however hexan-1-ol is not miscible with water. Explain this observation.

(2 marks) **KA4**

Total: 16 marks

2. Potassium chloride (KCl) is used to treat or prevent low levels of potassium (hypokalaemia). Potassium can be low as a result of disease, certain medications or prolonged illnesses with vomiting and diarrhoea.

(a) Solid potassium chloride is placed into water. The salt is dissociated and subsequently hydrated. Explain what is meant by this chemical statement.

(3 marks) **KA1**

(b) Write an ionic equation to represent the dissociation of potassium chloride in water.

(2 marks) **KA1**

(c) Describe what happens to the ionic lattice when placed in the polar solvent.

(3 marks) **KA4**

(d) When silver nitrate and potassium chloride are mixed together, a double-replacement reaction occurs. Explain what is meant by this statement.

(2 marks) **KA1**

(e) Explain how it is determined if a precipitate will form when two ionic solutions are placed together.

(2 marks) **KA1**

(f) Potassium chloride has also emerged to have water softening properties, therefore can be used to reduce hard water. Explain is meant by the term hard water.

(2 marks) **KA1**

(g) When hard water ions are present and soap scum forms, a magnesium or calcium carboxylate precipitate is formed. Describe what is the implication of this scum for its cleaning action.

(3 marks) **KA1**

Total: 17 marks

3. Magnesium carbonate ($MgCO_3$) is a white solid salt with a number of useful industrial applications including uses in flooring, fireproofing, cosmetics and even toothpastes.

(a) Calculate the number of moles of $MgCO_3$ contained in 30 mL of 0.300 mol L^{-1}.

(2 marks) **KA4**

(b) (i) Determine the concentration of a magnesium carbonate solution in mol L^{-1} prepared by dissolving 1.24 g of the solute into 200 mL of solution.

(4 marks) **KA4**

(ii) Convert this concentration to g L^{-1}

(3 marks) **KA4**

(iii) Convert this concentration to ppm

(2 marks) **KA1**

(c) Explain why grams per litre (g L^{-1}) is not the only unit used to represent mass concentration.

(2 marks) **KA1**

(d) Compare the terms primary standard and standard solution.

(2 marks) **KA1**

(e) Describe how mass–mass stoichiometry differs to that of mass–volume stoichiometry.

(2 marks) **KA1**

(f) Describe how to make 300 mL of 0.400 mol L^{-1} $MgCO_3$ solution from a stock solution of concentration 1.00 mol L^{-1}

(2 marks) **KA1**

Total: 13 marks

4. Potassium nitrate (KNO_3) is a soluble source of two essential plant nutrients. It is commonly used as a fertiliser for high-value crops that benefit from nitrate (NO_3^-) nutrition, and a source of potassium (K^+) free from chloride (Cl^-).

(a) When potassium nitrate dissolves in water, it has an enthalpy of solution of +35 kJ mol^{-1}. Write a thermochemical equation to represent this information.

(4 marks) **KA4**

(b) Compare the energy of the reactants and products in an endothermic reaction. Use this answer to write a simplified expression for ΔH.

(3 marks) **KA4**

(c) During endothermic reactions, heat energy is absorbed from the surroundings. Explain what is meant by the term 'surroundings'.

(2 marks) **KA1**

(d) Describe how the lattice and hydration energy of an ionic compound determines its enthalpy of solution.

(2 marks) **KA4**

Total: 14 marks

Topic 5: Acids and bases

5.1 Acid–base concepts

Science understanding

Acids are compounds or ions that donate protons, whereas bases are compounds or ions that accept protons, which are H^+ ions.

Acids are commonplace in our everyday lives from the citrus fruit that we eat to the soft drinks that we consume. Bases are less likely to be identified, however. These are found in various cleaning agents containing ammonia and in antacid tablets that are required for relieving heartburn or indigestion.

Science as a human endeavour:

The word **acid** is derived from the Latin word *acēre* meaning '*sour-tasting*'. The term **alkali** is derived from the Arabic word *al qalīy* meaning 'calcined ashes' was used to describe an alkaline solution containing the ashes of plant matter. An alkali substance is a water-soluble base, and the resulting solution is called an **alkaline solution**.

The first theoretical interpretation of an acid was provided by the French scientist Antoine Lavoisier in 1776. Lavoisier dissolved oxides of non-metals in water and determined that the resulting solution was acidic. Using Joseph Priestly's nitrous air test, Lavoisier knew that when a nitrogen oxide compound dissolved in water, the resulting solution was also acidic. Lavoisier concluded from these studies that the presence of oxygen within a compound determined the acidic properties of that substance. This was extremely important in the history of chemically characterising acids and bases, as this was the first of this nature. However, it was later proven to be incorrect.

In 1815, English chemist Sir Humphry Davy proposed that all acids contained hydrogen. However, it was known to scientists that molecular hydrogen and water contained hydrogen and neither compound was classified as acidic. In 1838, Justus Liebig offered a new proposal that an acid was a substance that contained one or more hydrogen atoms which could be replaced by metal atoms. This was neither correct nor complete, and required development from other scientific minds. Liebig did not define bases beyond saying that, when reacted with acids, they were able to neutralise the acid.

The first accurate theory of acids and bases was identified in the 1880s by Swedish chemist Svante Arrhenius. Arrhenius defined acids as compounds that produce hydrogen ions (H^+) in water. He defined bases as substances that produce hydroxide ions (OH^-) in water.

Based on these observations, Arrhenius defined an acid as a substance that, when added to water, increases the concentration of hydrogen ions (H^+) in the water. It is important to note that the atomic structure of a hydrogen cation is such that it has lost its one and only electron and thus is left with just a nucleus containing one proton. For this reason, a hydrogen cation is analogous to a proton.

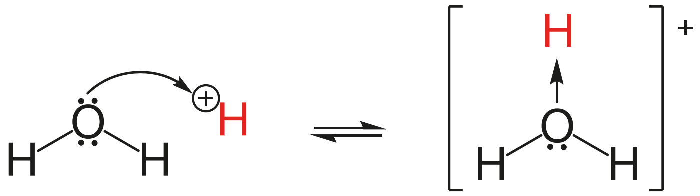
Figure 5.01: Hydrogen cation (proton)

Free protons are not present in solution. A single proton has a high electronegativity and accepts an electron pair from an oxygen atom in a molecule of water, forming a hydronium ion (H_3O^+). It is for this reason that the hydrogen ion and the hydronium ion are used interchangeably in acid–base chemistry. However, it is important to note that protons do not exist freely in solution.

$$H_2O_{(l)} + H^+_{(aq)} \rightarrow H_3O^+_{(aq)}$$

The hydrogen ion does not have an electron to contribute to the covalent bond and so the oxygen atom within a molecule of water contributes both electrons to the covalent bond formed with the hydrogen ion.

This forms a highly polarised covalent bond which is termed a **coordinate bond**. This occurs when both electrons shared between two non-metal atoms are contributed by just one of the bonding atoms. The structural formula is drawn with an arrow indicating the atom that is contributing both electrons to the covalent bond.



For example, when hydrochloric acid (HCl) is dissolved in sufficient water (H_2O), a proton is transferred from the acid to a molecule of water, which forms a hydronium ion (H_3O^+). This process is termed **ionisation**.

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$$

Arrhenius then defined a base as a substance that, when added to water, increases the concentration of hydroxide ions (OH^-) in solution. When ammonia (NH_3) is diluted with water (H_2O), molecules of water transfer a proton to ammonia to form ammonium hydroxide (NH_4OH). Ammonium hydroxide is water-soluble and ionises fully in solution forming hydroxide anions.

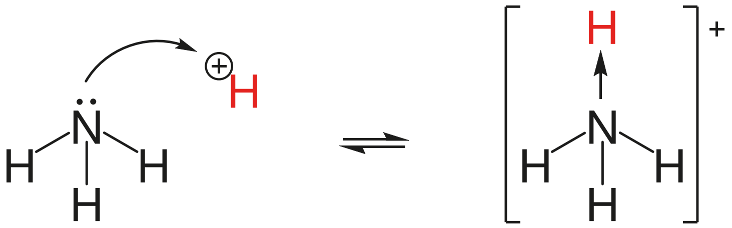
$$NH_{3(aq)} + H_2O_{(l)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)}$$

Arrhenius's theory of acids and bases is limited to the reaction of acids and bases when diluted with water.

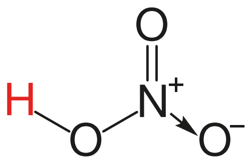
Science as a human endeavour:

In 1923, Danish scientist Johannes Nicolaus Brønsted and English chemist Thomas Martin Lowry independently developed a new theory on acids and bases. Brønsted and Lowry extended Arrhenius's theory and defined acids and bases in terms of **proton transfer**.

Acids donate protons and bases accept protons. This new theory was able to describe how a base like ammonia was able to increase the concentration of hydroxide ions in solution when reacted with water. Ammonia (a base) accepts a proton from a molecule of water to produce the ammonium ion and a hydroxide ion.



Another molecular compound exhibiting a coordinate bond is nitric acid (HNO_3). The nitrogen atom contributes both electrons to the covalent bond formed with the hydrogen ion.



Nitric acid is a compound that is defined as an acid due to its ability to donate a proton to a base in a chemical reaction. Water (H_2O) accepts a proton when reacted with nitric acid (HNO_3); this forms hydronium ions (H_3O^+).

$$HNO_{3(aq)} + H_2O_{(l)} \rightarrow NO^-_{3(aq)} + H_3O^+_{(aq)}$$

Brønsted and Lowry's theory states that a base accepts a proton from an acid in a chemical reaction. A molecule of ammonia accepts a proton from an acid to form an ammonium ion.

$$NH_{3(aq)} + H^+_{(aq)} \rightarrow NH^+_{4(aq)}$$

In the reactions above, water is shown to accept and donate protons in a chemical reaction. Water is an example of an **amphiprotic substance.** An amphiprotic substance is a substance that can both accept and donate protons.

Acid–base equations

Science understanding

The reactions between acids and bases can be represented using chemical equations that illustrate the transfer of protons.

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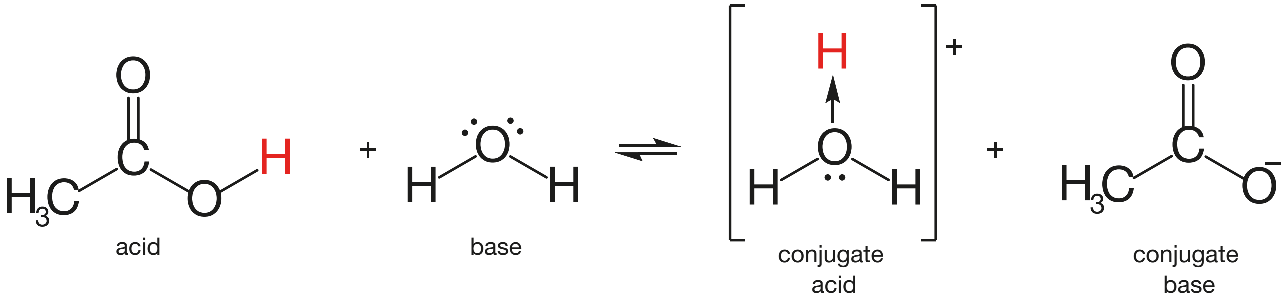
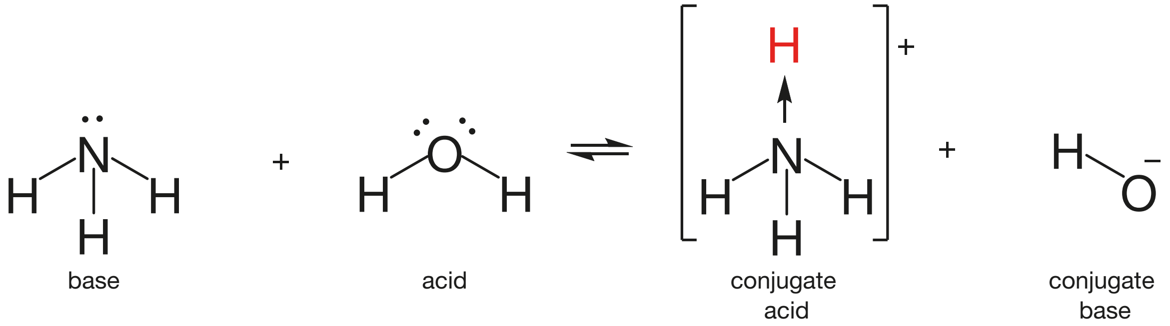
Ethanoic acid (CH_3COOH) is an acid which partially ionises in water (H_2O) yielding hydronium (H_3O^+) and ethanoate (CH_3COO^-) ions.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

Ethanoate ions are able to accept a proton from hydronium ions and, therefore, the reaction is reversible. The hydronium ions are termed the **conjugate acid** and ethanoate ions the **conjugate base**.



When ammonia reacts with water, a molecule of ammonia accepts a proton from a molecule of water to form ammonium and hydroxide ions. The ammonium ion can transfer a proton to the hydroxide ion and, therefore, the reaction is reversible. The ammonium ion (NH_4^+) is the conjugate acid and the hydroxide ion is the conjugate base.



Questions

1. Nitric acid (HNO_3) is classified as an acid. Define an acid, and explain your answer by using a balanced equation involving nitric acid reacting with water.

(3 marks) **KA1**

2. Sodium hydroxide (NaOH) is classified as a base. Define a base, and explain your answer by using a balanced equation.

(3 marks) **KA1**

3. Identify the acid, base, conjugate acid and conjugate base in the reactions below.

   (a) $CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

   ………… ………… ………… …………

   (b) $CO_{3\,(aq)}^{2-} + H_2O_{(l)} \rightleftharpoons HCO^-_{3(aq)} + OH^-_{(aq)}$

   ………… ………… ………… …………

   (c) $H_2SO_{4(aq)} + NH_{3(aq)} \rightarrow HSO^-_{4(aq)} + NH^+_{4(aq)}$

   ………… ………… ………… …………

   (d) $KOH_{(aq)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(aq)}$

   ………… ………… ………… …………

(8 marks) **KA4**

4. Write balanced chemical equations for the following reactions and clearly indicate the conjugate acid–base pairings.

   (a) hydrochloric acid and water

   ……………………………………………………

   (b) sodium hydroxide and hydrochloric acid

   ……………………………………………………

   (c) ethanoic acid and ammonia

   ……………………………………………………

   (d) ammonium ion and water

   ……………………………………………………

(16 marks) **KA4**

Strength of acids and bases

Science understanding

The strength of acids is explained by the degree of ionisation in aqueous solution.

Acidic and basic compounds can be further classified according to their strength. Acids and bases ionise in water, and it is the degree of ionisation that occurs in the reaction that defines the strength of acids and bases. A strong acid will completely ionise in water such that one mole of acid (reactant) yields one mole of hydronium ions (product). Hydrochloric acid (HCl) will completely ionise in water (H_2O) to produce chloride (Cl^-) and hydronium (H_3O^+) ions (refer to Figure 5.02). Hydrochloric acid is termed a strong acid, while its conjugate base (the chloride ion) is a weak base.

$$HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$$

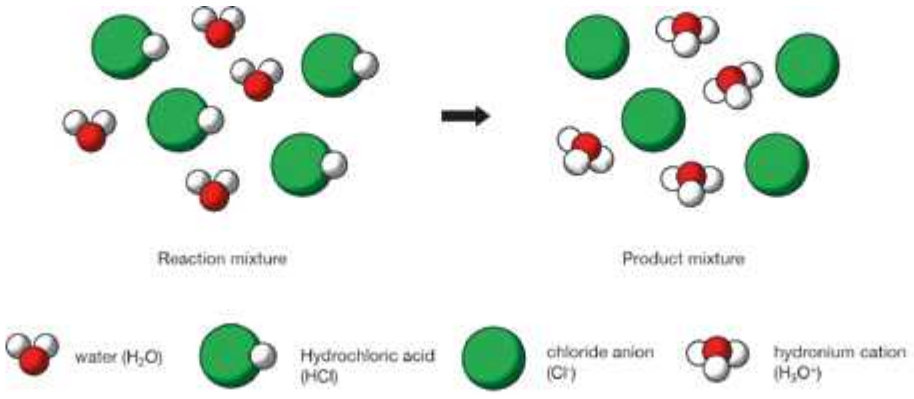


Figure 5.02: Complete ionisation of hydrochloric acid

Conversely, a weak acid partially ionises in water such that one mole of a weak acid produces less than one mole of hydronium ions. Ethanoic acid (CH_3COOH) is a weak acid, as when it is reacted with water (H_2O), very few moles of ethanoic acid are ionised to form hydronium (H_3O^+) and ethanoate (CH_3COO^-) ions. Therefore, the solution will have a greater proportion of ethanoic acid molecules to ethanoate ions (refer to Figure 5.03).

$$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

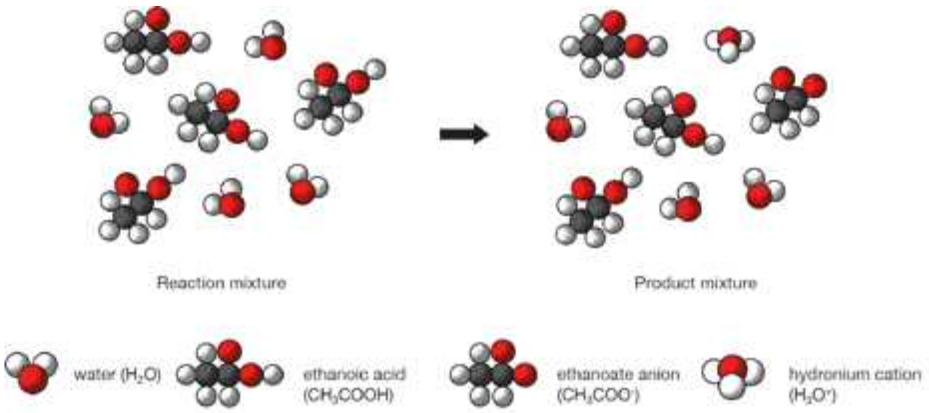


Figure 5.03: Partial ionisation of ethanoic acid

Strong bases ionise completely in water such that one mole of base yields one mole of hydroxide ions in solution. Sodium hydroxide (NaOH) is a strong acid, as it will completely ionise in water to produce sodium (Na^+) and hydroxide (OH^-) ions (refer to Figure 5.04).

$$NaOH_{(s)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

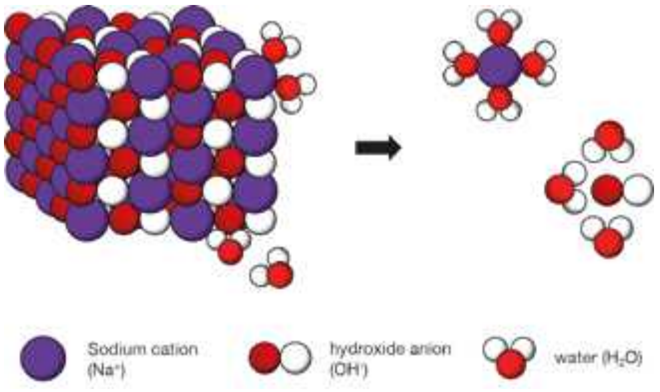


Figure 5.04: Complete ionisation of sodium hydroxide

Conversely, a weak base partially ionises in water such that one mole of base yields less than one mole of hydroxide ions. Ammonia (NH_3) is a weak base and will only partially ionise in water (H_2O) to produce ammonium (NH_4^+) and hydroxide (OH^-) ions (refer to Figure 5.05).

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

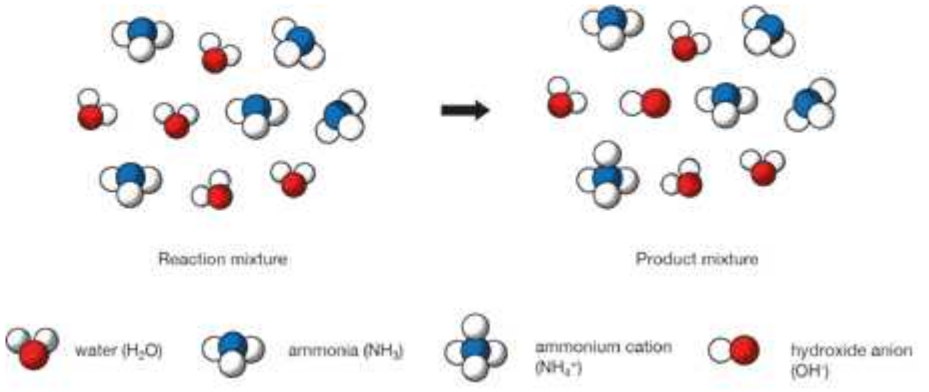


Figure 5.05: Partial ionisation of ammonia

Acid strength and **acid concentration** are two terms which have different meanings in chemistry. Acid strength refers to the degree of ionisation a substance undergoes when it is reacted with water. Acid concentration refers to the number of moles of the acid per litre of solution.

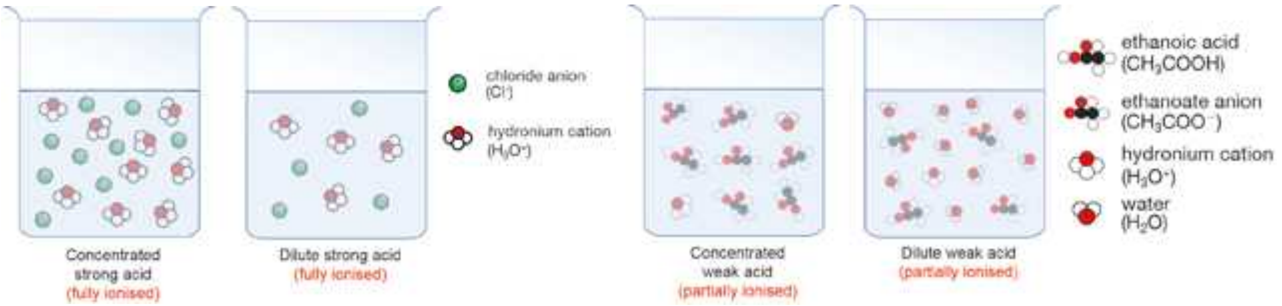


Figure 5.06: Acid strength vs acid concentration

Hydrofluoric acid has a molar concentration of 28 mol L^{-1} but only partially ionises in solution.

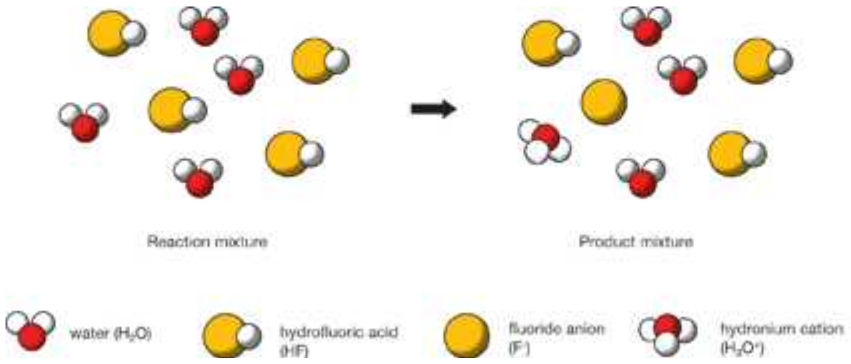


Figure 5.07: Partial ionisation of hydrofluoric acid

Strong acids	Weak acids	Weak bases	Strong bases
sulfuric acid, H_2SO_4	sulfurous acid, H_2SO_3	ammonia, NH_3	sodium hydroxide, NaOH
nitric acid, HNO_3	carbonic acid, H_2CO_3	sodium carbonate, Na_2CO_3	potassium hydroxide, KOH
hydrochloric acid, HCl	phosphoric acid, H_3PO_4	potassium hydrogen carbonate, $KHCO_3$	
	ethanoic acid, CH_3COOH		

Table 5.01: Common acids and bases

Questions

5. Sulfuric acid (H_2SO_4) is classified as a strong acid, whereas carbonic acid (H_2CO_3) is classified as a weak acid. Explain the difference between strong and weak acids, and support your answer by using balanced equations involving these two acids.

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....*(6 marks)* **KA1**

Polyprotic acids

Science understanding

> Acids can be described as monoprotic or polyprotic, depending on the number of protons available for donation.
>
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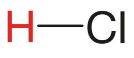
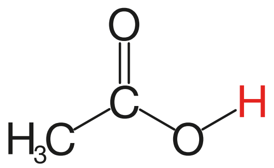
An acid that transfers one proton to a base in a chemical reaction is termed a **monoprotic acid**. Examples of monoprotic acids are hydrochloric acid (HCl) and ethanoic acid (CH_3COOH).

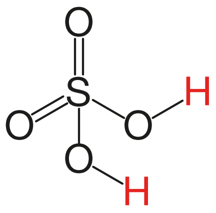


hydrochloric acid

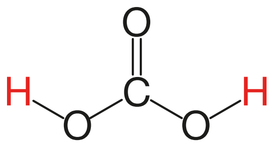


ethanoic acid

An acid that can transfer more than one proton to a base in a chemical reaction is termed a **polyprotic acid**. Acids that transfer up to two protons are termed **diprotic acids** and examples of these acids are sulfuric acid (H_2SO_4) and carbonic acid (H_2CO_3).



sulfuric acid



carbonic acid

The ionisation of a diprotic acid occurs in two stages. Consider sulfuric acid (H_2SO_4) ionising in water (H_2O). In the first stage, one proton is transferred to water to form hydronium (H_3O^+) and hydrogen sulfate (HSO_4^-) ions.

$$H_2SO_{4(aq)} + H_2O_{(l)} \rightarrow HSO^-_{4(aq)} + H_3O^+_{(aq)}$$

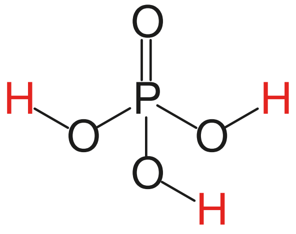
The hydrogen sulfate ion (HSO_4^-) then transfers a proton to a second molecule of water (H_2O) to form sulfate (SO_4^{2-}) and hydronium (H_3O^+) ions. Hydrogen sulfate is a weak acid and therefore there will still be a large proportion of hydrogen sulfate ions in solution.

$$HSO^-_{4(aq)} + H_2O_{(l)} \rightleftharpoons SO^{2-}_{4(aq)} + H_3O^+_{(aq)}$$

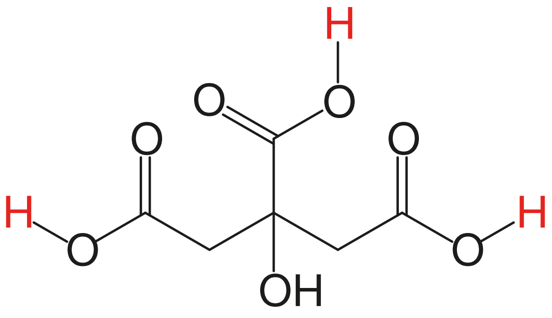
This series of reactions is summarised in a single balanced equation.

$$H_2SO_{4(aq)} + 2H_2O_{(l)} \rightarrow SO^{2-}_{4(aq)} + 2H_3O^+_{(aq)}$$

Acids which transfer up to three protons are termed **triprotic acids** and examples of these acids are phosphoric acid (H_3PO_4) and citric acid ($C_6H_8O_7$).



phosphoric acid



citric acid

The ionisation of a triprotic acid occurs in three stages. Consider phosphoric acid (H_3PO_4) ionising in water (H_2O). In the first stage, one proton is transferred to water to form hydronium (H_3O^+) and dihydrogen phosphate ($H_2PO_4^-$) ions.

$$H_3PO_{4(aq)} + H_2O_{(l)} \rightarrow H_2PO^-_{4(aq)} + H_3O^+_{(aq)}$$

The dihydrogen phosphate ($H_2PO_4^-$) ion then transfers a proton to a second molecule of water (H_2O) to form the monohydrogen phosphate (HPO_4^{2-}) and hydronium (H_3O^+) ions. Dihydrogen phosphate is a weak acid and therefore will not completely ionise in water.

$$H_2PO^-_{4(aq)} + H_2O_{(l)} \rightleftharpoons HPO^{2-}_{4(aq)} + H_3O^+_{(aq)}$$

The monohydrogen phosphate (HPO_4^{2-}) ion then transfers a proton to a third molecule of water (H_2O) to form phosphate (PO_4^{3-}) and hydronium (H_3O^+) ions. Monohydrogen phosphate is a weak acid and, therefore, when reacted with water, will only produce a small proportion of phosphate ions. This is due to monohydrogen phosphate being a highly charged ion; it is therefore more likely to accept protons. This means it is an amphiprotic substance.

$$HPO^{2-}_{4(aq)} + H_2O_{(l)} \rightleftharpoons PO^{3-}_{4(aq)} + H_3O^+_{(aq)}$$

This series of reactions is summarised in a single balanced equation.

$$H_3PO_{4(aq)} + 3H_2O_{(l)} \rightleftharpoons PO^{3-}_{4(aq)} + 3H_3O^+_{(aq)}$$

Questions

6. An ammonium ion is classified as a monoprotic acid. Explain what this means, and use a chemical equation to support your answer.

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.... *(3 marks)* **KA1**

7. Classify hydrogen sulfide as a mono-, di- or triprotic acid and write balanced chemical equations supporting your answer.

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.... *(5 marks)* **KA1**

8. Show the ionisation of boric acid (H_3BO_4), a triprotic acid, via a series of balanced chemical equations.

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.... *(6 marks)* **KA4**

Acid–base indicators

Science understanding

> Acid–base indicators are weak acids or bases where the acidic form is of a different colour from the basic form.
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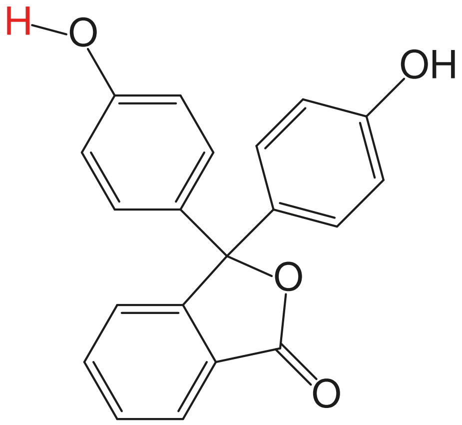
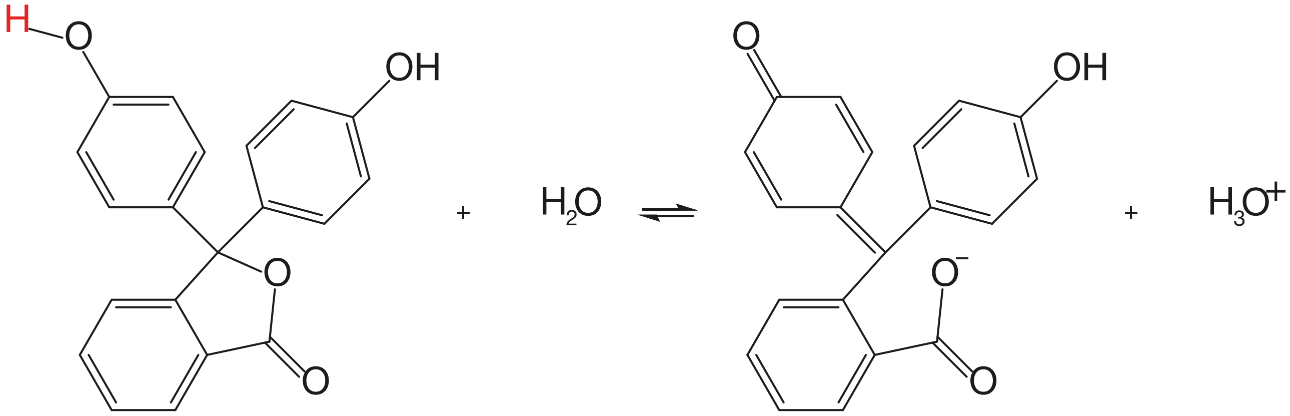
Acid-base indicators can be used to determine whether a solution is acidic, basic or neutral, because indicators change colour to provide a visual representation of a solution's pH level. Indicators are substances comprised of molecules of a weak acid and its conjugate-base pair. Both substances are a different colour depending on the concentration of hydronium ions (H_3O^+) in the solution. An example of this is a molecule called phenolphthalein.



Phenolphthalein is colourless in an acidic solution. When placed in a basic solution, a proton is transferred from the indicator to the base to form an anion that reflects red and violet wavelengths of light. This gives the indicator a pink colour.



acid base

The reason phenolphthalein solution is colourless is because the molecule does not emit visible wavelengths of light. The chemical formula for phenolphthalein is simplified to HIn.

The following equation can be used to show the reaction of phenolphthalein in a basic solution .

$$HIn + OH^- \longrightarrow In^- + H_2O$$

colourless → deep pink

If the phenolphthalein ion (In^-) is placed in an acidic solution the following reaction occurs.

$$In^- + H^+ \longrightarrow HIn$$

deep pink → colourless

All acid–base indicators reflect different wavelengths (colours) of visible light depending on the pH of the solution.

Indicator	Colour in acid (pH < 7)	Colour at pH = 7	Colour in base (pH > 7)
Phenolphthalein	**colourless**	**colourless**	**pink, red**
Bromothymol blue	**yellow**	**green**	**blue**
Red litmus	**red**	**red**	**blue**
Blue litmus	**red**	**blue**	**blue**
Universal indicator	**red, orange, yellow**	**green**	**Blue, violet, purple**

Table 5.02: Common indicators used in the laboratory

Questions

9. Carbonic acid is classified as a diprotic acid. Define a diprotic acid and write balanced equations to support your answer.

(5 marks) **KA1**

10. State the acidic or basic character of the following substances based on their reactions with the given indicators.

    (a) Compound A is placed in a solution containing phenolphthalein and the solution changes from colourless to pink.

    (b) A sample of solution X is placed onto blue litmus paper and it turns the litmus a red colour.

    (c) A student collects two samples of a colourless solution C and places them into separate test tubes. In one test tube the student places a few drops of phenolphthalein and the solution remains colourless. In the second test tube the student places a few drops of universal indicator and the solution turns green.

(3 marks) **KA1**

11. Sulfur is able to attract two electrons to become the sulfide ion.

    (a) Write the electronic configuration for a sulfide ion.

(2 marks) **KA4**

    (b) Sulfide ions are classified as basic. Using a balanced chemical equation describing the reaction of sulfide ions with water, explain why sulfide ions are basic.

(3 marks) **KA1**

    (c) State the colour that red litmus becomes when it is placed in a sample of the resulting solution.

(1 mark) **KA1**

5.2: Reactions of acids and bases

When positioning elements on the periodic table, scientists identified similar properties that elements had and grouped them accordingly. Acids and bases are two broad categories which can also be defined according to the similar chemical reactions that they experience. Therefore, acidic and basic properties also help to determine the position of elements on the periodic table.

Acidic oxides

Science understanding

The oxides of non-metals are commonly acidic and generate oxyacids when dissolved in water.

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Oxides that react with water to form acids are called **acidic oxides**. The oxides of non-metallic elements react chemically with water to form acids. Therefore, the oxides of non-metallic elements are acidic oxides. An example of an acidic oxide is sulfur dioxide (SO_2). Sulfur dioxide reacts with water (H_2O) to form sulfurous acid (H_2SO_3), according to the following equation.

$$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

When non-metallic oxides react with bases, they form a salt as well as water. Sulfur dioxide reacts with sodium hydroxide (NaOH), a base, to form sodium sulfite (Na_2SO_3) and water (H_2O).

$$SO_{2(g)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{3(aq)} + H_2O_{(l)}$$

Using the balanced chemical equation above, we can write a net ionic equation using the steps covered in Topic 4.2. Sodium hydroxide (NaOH) ionises in solution to produce sodium (Na^+) and hydroxide (OH^-) ions. Sodium sulfite (Na_2SO_3) also ionises in solution to produce sodium (Na^+) and sulfite (SO_3^{2-}) ions. This chemical equation can be written as a complete ionic equation.

$$SO_{2(g)} + 2Na^+_{(aq)} + 2OH^-_{(aq)} \rightarrow 2Na^+_{(aq)}\ SO^{2-}_{3(aq)} + H_2O_{(l)}$$

Sodium ions (Na^+) are found as a reactant and a product in the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$SO_{2(g)} + 2OH^-_{(aq)} \rightarrow SO^{2-}_{3(aq)} + H_2O_{(l)}$$

The sulfite anion (SO_3^{2-}) is an example of an **oxyanion**. Thus, when acidic oxides react with bases an oxyanion and water are produced. All other non-metal oxides will react in this way and their products are summarised in table 5.03.

Note: Ensure that when writing the net ionic equation for acid–base reactions that special attention is given to the solubility rules (Topic 4.2).

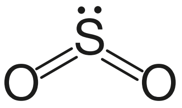
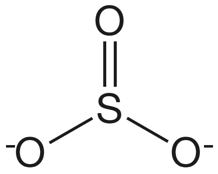
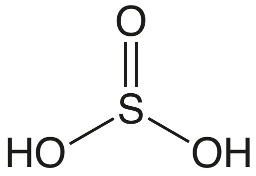
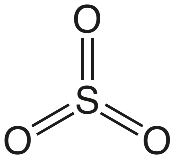
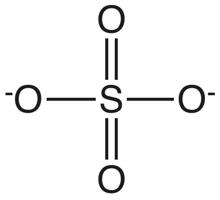
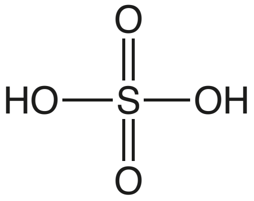
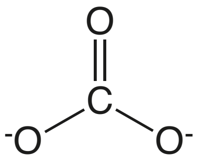
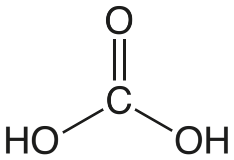
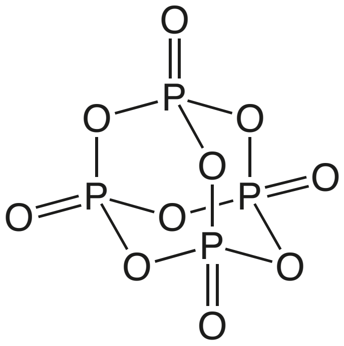
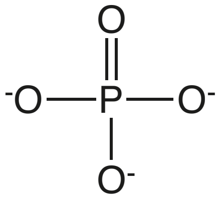
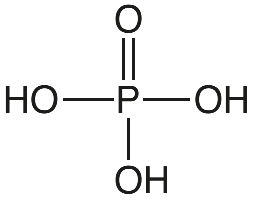
Non-metal oxide	Oxyanion	Acid
 sulfur dioxide, SO_2	 sulfite, SO_3^{2-}	 sulfurous acid, H_2SO_3
 sulfur trioxide, SO_3	 sulfate, SO_4^{2-}	 sulfuric acid, H_2SO_4
$O{=}C{=}O$ carbon dioxide, CO_2	 carbonate, CO_3^{2-}	 carbonic acid, H_2CO_3
 phosphorous pentoxide, P_4O_{10}	 phosphate, PO_4^{3-}	 phosphoric acid, H_3PO_4
nitrogen dioxide, NO_2	nitrate, NO_3^-	nitric acid, HNO_3
dichlorine monoxide, Cl_2O	hypochlorite, ClO^-	hypochlorous acid, HOCl

Figure 5.09: Common non–metal oxides

Note: Structural formulae are drawn for those molecules which are required to be drawn at Stage 1

Questions

12. Write balanced chemical equations for the following oxides when reacted with water.

    (a) sulfur trioxide

    (b) carbon dioxide

    (c) dichlorine monoxide

(6 marks) **KA4**

13. Write fully balanced chemical equations and net ionic equations for the following oxides when reacted with sodium hydroxide.

    (a) sulfur dioxide

    Fully balanced chemical equation:

    Net ionic equation:

    (b) phosphorous pentoxide

    Fully balanced chemical equation:

    Net ionic equation:

    (c) nitrogen dichlorine monoxide

    Fully balanced chemical equation:

    Net ionic equation:

(12 marks) **KA4**

14. Draw structural formulae for the following compounds.

    (a) sulfur trioxide

    (b) carbon dioxide

    (c) sulfuric acid

    (d) phosphoric acid

(8 marks) **KA4**

Basic oxides

Science understanding

Metal oxides are commonly basic. *© Copyright SACE 2022*

Oxides that react with water to form hydroxide ions are called **basic oxides**. The oxides of metals react chemically with molecules of water to form hydroxide ions (OH^-). Therefore, metallic oxides are basic oxides. One such basic oxide is magnesium oxide. Magnesium oxide (MgO) dissolves partially in water (H_2O) forming magnesium (Mg^{2+}) and hydroxide (OH^-) ions according to the following equation.

$$MgO_{(s)} + H_2O_{(l)} \rightarrow Mg^{2+}_{(aq)} + 2OH^-_{(aq)}$$

Metallic oxides will also react with acidic substances to form a cation and water. Magnesium oxide (MgO) reacts chemically with a solution of hydrochloric acid (HCl) to form magnesium chloride ($MgCl_2$) and water (H_2O).

$$MgO_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_2O_{(l)}$$

Using the balanced chemical equation above we can write a net ionic equation. Hydrochloric acid (HCl) completely ionises in solution to produce hydrogen (H^+) and chloride (Cl^-) ions. Magnesium chloride ($MgCl_2$) also ionises in solution and produces magnesium (Mg^{2+}) and chloride (Cl^-) ions. This chemical equation can be written as a complete ionic equation.

$$MgO_{(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)} + H_2O_{(l)}$$

Chloride ions (Cl^-) are found as both a reactant and a product in the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$MgO_{(s)} + 2H^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + H_2O_{(l)}$$

Questions

15. Write balanced chemical equations for the following oxides when reacted with water.

 (a) potassium oxide

 ....

 (b) calcium oxide

 ....

 (c) sodium oxide

 .... *(6 marks)* **KA4**

16. Write fully balanced chemical equations and net ionic equations for the following oxides when reacted with sulfuric acid.

 (a) sodium oxide
 Fully balanced chemical equation:

 ....

 Net ionic equation:

 ....

 (b) copper(II) oxide
 Fully balanced chemical equation:

 ....

 Net ionic equation:

 ....

strontium oxide

Fully balanced chemical equation:

.. ..

Net ionic equation:

. ...*(12 marks)* **KA4**

Amphoteric oxides

A number of elemental oxides exhibit properties characteristic of both acids and bases. For example, lead oxide neutralises acids and bases. Oxides that can react as both an acid and a base are called **amphoteric**, *a term which comes* from the Greek word *amphoteros* meaning "both".

Aluminium oxide (Al_2O_3) can neutralise hydrochloric acid (HCl) according to the following balanced chemical equation.

$$Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_2O_{(l)}$$

Using the balanced chemical equation above, we can write a net ionic equation. Hydrochloric acid (HCl) completely ionises in solution to produce hydrogen (H^+) and chloride (Cl^-) ions. Aluminium chloride ($AlCl_3$) also ionises in solution and produces aluminium (Al^{3+}) and chloride (Cl^-) ions. This chemical equation can be written as a complete ionic equation.

$$Al_2O_{3(s)} + 6H^+_{(aq)} + 6Cl^-_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 6Cl^-_{(aq)} + 3H_2O_{(l)}$$

Chloride ions (Cl^-) are found as reactants and products in this reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$Al_2O_{3(s)} + 6H^+_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 3H_2O_{(l)}$$

Conversely, aluminium oxide (Al_2O_3) neutralises sodium hydroxide (NaOH) solution to form sodium aluminate ($NaAlO_2$) and water (H_2O).

$$Al_2O_{3(s)} + 2NaOH_{(aq)} \longrightarrow 2NaAlO_{2(aq)} + H_2O_{(l)}$$

Sodium hydroxide (NaOH) completely ionises in solution to produce sodium (Na^+) and hydroxide (OH^-) ions. Sodium aluminate ($NaAlO_2$) also ionises in solution and produces sodium (Na^+) and aluminate (AlO_2^-) ions. This chemical equation can be written as a complete ionic equation.

$$Al_2O_{3(s)} + 2Na^+_{(aq)} + 2OH^-_{(aq)} \longrightarrow 2Na^+_{(aq)} + 2AlO^-_{2(aq)} + H_2O_{(l)}$$

Sodium ions (Na^+) are found as reactants and products in this reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$Al_2O_{3(s)} + 2OH^-_{(aq)} \longrightarrow 2AlO^-_{2(aq)} + H_2O_{(l)}$$

Therefore, aluminium oxide is an amphoteric oxide as it exhibits properties characteristic of both acid and base.

Metalloid oxide	Oxyanion
aluminium oxide, Al_2O_3	aluminate, AlO_2^-
zinc oxide, ZnO	zincate, ZnO_2^{2-}

Table 5.03: common amphoteric oxides and the oxyanions that they form

Questions

17. Zinc oxide is classified as an amphoteric oxide.

    (a) Define an amphoteric oxide.

    (2 marks) **KA1**

    (b) Write balanced ionic equations to describe the amphoteric nature of zinc oxide.

    (i)

    (ii)

    (4 marks) **KA4**

18. State whether the acidic nature of elements increases or decreases as you move across the periodic table.

    (1 mark) **KA1**

19. State the acid/base nature of the following oxides.

    (a) potassium

    (b) aluminium

    (c) carbon

    (3 marks) **KA4**

Neutralisation reactions

Science understanding

Similarities in the reactions of different acids with bases (metal oxides, hydroxides, and carbonates) allow products to be predicted from known reactants.

Neutralisation is an exothermic reaction.

An acid–base reaction is classified simply as a neutralisation reaction. This is because if one mole of a strong acidic compound was to react with exactly one mole of a strong basic compound the resulting solution would be neutral.

When aqueous solutions of acids and bases are mixed, the hydrogen ions (H^+) in the acid chemically combine with the hydroxide ions (OH^-) from the base to form molecules of water (H_2O). The reaction is exothermic (bonds formed) and is described by the following thermochemical equation.

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)} \quad \Delta H = -55.9 \text{ kJ mol}^{-1}$$

55.9 kJ of energy are released when one mole of hydrogen and hydroxide ions chemically combine to form one mole of water. Therefore, neutralisation reactions between strong acids and bases are exothermic reactions and subsequently possess a negative enthalpy value.

Acids reacting with metal oxides

Acids react with metallic oxides (base) to form a salt and water.

$$\text{metal oxide} + \text{acid} \rightarrow \text{salt} + \text{water}$$

When sodium oxide (Na_2O) reacts with sulfuric acid (H_2SO_4), the resulting solution is made up of water (H_2O) and sodium sulfate (Na_2SO_4). However, if you were to evaporate off the liquid water then the sodium and sulfate ions would crystallise and form a solid salt.

$$Na_2O_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$$

In this example, the sodium sulfate (Na_2SO_4) is dissolved in water (H_2O) and therefore exists as sodium (Na^+) and sulfate (SO_4^{2-}) ions. Sulfuric acid (H_2SO_4) is a strong acid and so completely ionises in solution to produce hydrogen (H^+) and sulfate (SO_4^{2-}) ions.

$$Na_2O_{(s)} + 2H^+_{(aq)} + SO^{2-}_{4(aq)} \rightarrow 2Na^+_{(aq)} + SO^{2-}_{4(aq)} + H_2O_{(l)}$$

Sulfate ions (SO_4^{2-}) are found as both a reactant and a product in the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$Na_2O_{(s)} + 2H^+_{(aq)} \rightarrow 2Na^+_{(aq)} + H_2O_{(l)}$$

Acids reacting with metal hydroxides

Acids react with metal hydroxides to form a salt and water.

$$\text{metal hydroxide} + \text{acid} \rightarrow \text{salt} + \text{water}$$

A solution of potassium hydroxide (KOH) will react with hydrochloric acid (HCl) and form potassium chloride (KCl) and water (H_2O).

$$KOH_{(aq)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}$$

In this example, potassium hydroxide (KOH), hydrochloric acid (HCl) and potassium chloride (KCl) all ionise in solution. Therefore, the complete ionic equation representing this reaction can be written as follows.

$$K^+_{(aq)} + OH^-_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)} \rightarrow K^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)}$$

Potassium (K^+) and chloride (Cl^-) ions are found as reactants and products in the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(l)}$$

This net ionic equation can be used for all soluble hydroxides. However, not all hydroxides are soluble (see Topic 4.2) and, consequently, these reactions and equations need to be considered. Calcium hydroxide ($Ca(OH)_2$), an insoluble hydroxide, reacts with hydrochloric acid (HCl) to form calcium chloride ($CaCl_2$) and water (H_2O).

$$Ca(OH)_{2(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$$

When you write the complete ionic equation, remember that calcium hydroxide cannot be separated into its ions as it does not ionise in water (it is insoluble). Hydrochloric acid (HCl) and calcium chloride ($CaCl_2$) both ionise in water. Therefore, the complete ionic equation representing this reaction can be written as follows.

$$Ca(OH)_{2(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow Ca^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2H_2O_{(l)}$$

Chloride ions (Cl^-) are found as reactants and products in the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$Ca(OH)_{2(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + 2H_2O_{(l)}$$

A solution of sodium hydroxide (NaOH), a strong base, will react with a solution of ethanoic acid (CH_3COOH), a weak acid, to form sodium ethanoate ($NaCH_3COO$) and water (H_2O).

$$NaOH_{(aq)} + CH_3COOH_{(aq)} \rightarrow NaCH_3COO_{(aq)} + H_2O_{(l)}$$

When you write complete ionic equations, remember that the strength of the acids and bases involved in the reaction need to be considered. Ethanoic acid is a weak acid and only partially ionises in water. Therefore, when you write the ionic equation, ethanoic acid cannot be separated into its ions. Sodium hydroxide (NaOH) and sodium ethanoate ($NaCH_3COO$) both ionise in water. Therefore, the complete ionic equation representing this reaction can be written as follows.

$$Na^+_{(aq)} + OH^-_{(aq)} + CH_3COOH_{(aq)} \rightarrow Na^+_{(aq)} + CH_3COO^-_{(aq)} + H_2O_{(l)}$$

Sodium ions (Na^+) are found as both a reactant and a product of the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$OH^-_{(aq)} + CH_3COOH_{(aq)} \rightarrow CH_3COO^-_{(aq)} + H_2O_{(l)}$$

Acids reacting with metal carbonates

Acids react with metal carbonates to form a salt and water, as well as effervescence (bubbles) in the form of carbon dioxide gas.

metal carbonate + acid → salt + water + carbon dioxide

Calcium carbonate ($CaCO_3$) dissolves in sulfuric acid (H_2SO_4) to form calcium sulfate ($CaSO_4$), which precipitates from solution.

$$CaCO_{3(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + CO_{2(g)} + H_2O_{(l)}$$

Sulfuric acid (H_2SO_4) is a strong acid and completely ionises in solution. Therefore, the complete ionic equation representing this reaction can be written as follows.

$$CaCO_{3(s)} + 2H^+_{(aq)} + SO^{2-}_{4(aq)} \rightarrow CaSO_{4(s)} + CO_{2(g)} + H_2O_{(l)}$$

There are no spectator ions in this reaction and therefore the complete ionic equation is also the net ionic equation.

Calcium carbonate ($CaCO_3$) dissolves in nitric acid (HNO_3) forming calcium nitrate ($Ca(NO_3)_2$), water (H_2O) and carbon dioxide gas (CO_2).

$$CaCO_{3(s)} + 2HNO_{3(aq)} \rightarrow Ca(NO_3)_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

In this example, nitric acid (HNO_3) and calcium nitrate ($Ca(NO_3)_2$) both ionise in solution. Therefore, the complete ionic equation representing this reaction can be written as follows.

$$CaCO_{3(s)} + 2H^+_{(aq)} + 2NO^-_{3(aq)} \rightarrow Ca^{2+}_{(aq)} + 2NO^-_{3(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Nitrate ions (NO_3^-) are found as reactants and products of the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Acids reacting with metal hydrogencarbonates (bicarbonates)

Acids react with metal hydrogencarbonates to form a salt and water, as well as effervescence (bubbles) in the form of carbon dioxide gas.

metal hydrogencarbonates + acid → salt + water + carbon dioxide

Carbon dioxide gas (CO_2) is formed from the reaction of sodium hydrogencarbonate ($NaHCO_3$) and hydrochloric acid (HCl) solution.

$$NaHCO_{3(s)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

In this example, hydrochloric acid (HCl) and sodium chloride ($NaCl$) both ionise in solution. Therefore, the complete ionic equation representing this reaction can be written as follows.

$$NaHCO_{3(s)} + H^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Chloride ions (Cl^-) are found as both a reactant and a product of the reaction, and therefore are spectator ions. The complete ionic equation can be summarised as a net ionic equation as follows.

$$NaHCO_{3(s)} + H^+_{(aq)} \rightarrow Na^+_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Questions

20. Write fully balanced chemical equations to describe the following acid–base reactions.

    (a) zinc oxide and sulfuric acid

....

    (b) aluminium hydroxide and hydrochloric acid

....

    (c) hydrogencarbonates and nitric acid

....

    (d) potassium carbonate and hydrochloric acid

....

(8 marks) **KA4**

21. Write fully balanced chemical equations and net ionic equations for the following neutralisation reactions.

    (a) Potassium hydroxide solution and nitric acid solution

    Fully balanced chemical equation:

....

    Net ionic equation:

....

    (b) Solid sodium carbonate and sulfuric acid solution

    Fully balanced chemical equation:

....

    Net ionic equation:

....

    (c) Solid potassium hydrogen carbonate and hydrochloric acid solution

    Fully balanced chemical equation:

....

    Net ionic equation:

....

(12 marks) **KA4**

Stoichiometric calculations involving neutralisation reactions

In chemical laboratories, neutralisation reactions are used to determine the concentration of acidic and basic solutions. This is achieved by making a standard solution of known concentration (refer to Topic 4.3.2) and reacting it with a second solution of unknown concentration. The volume of the standard solution required to neutralise the second solution is recorded. Stoichiometric calculations are then used to calculate the concentration of the second solution.

For example, when sodium hydroxide (NaOH) is reacted with hydrochloric acid (HCl) the following reaction occurs:

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

The equation indicates that the stoichiometric ratio is 1:1 and that one mole of sodium hydroxide will react with exactly one mole of hydrochloric acid to form one mole of sodium chloride (NaCl) and one mole of water (H_2O). The number of moles of acid required to neutralise a given number of moles of base can be predicted using stoichiometry.

Consider the following example: 0.5012 g of sodium hydroxide was dissolved in 100 mL of water to form a solution. The solution was then neutralised completely by reaction with a 100 mL solution of hydrochloric acid. Determine the molar concentration of hydrochloric acid solution in this reaction.

1. Find the number of moles of sodium hydroxide reacting with the hydrochloric acid.

   $m_{(NaOH)} = 0.5012 \text{ g}$

   $M_{(NaOH)} = (M_{Na} \times 1) + (M_O \times 1) + (M_H \times 1)$

   $M_{(NaOH)} = (22.99) + (16.00) + (1.008)$

   $M_{(NaOH)} = 39.998 \text{ g mol}^{-1}$

   $n_{(NaOH)} = \dfrac{m_{(NaOH)}}{M_{(NaOH)}}$

   $n_{(NaOH)} = 0.5012$

   $n_{(NaOH)} = 0.01253 \text{ mol}$

2. Using the mole ratio, find the number of moles of hydrochloric acid.

   From the reaction equation we can see that the mole ratio is:

   $\underline{NaOH} : \underline{HCl}$
   $1 : 1$

   Therefore, the number of moles of sodium hydroxide is equal to the number of moles of hydrochloric acid.

   $n_{(HCl)} = 0.01253 \text{ mol}$

3. Calculate the concentration of hydrochloric acid.

   $v_{(HCl)} = 100 \text{ mL}$

   $v_{(HCl)} = 0.100 \text{ L}$

   $c_{(HCl)} = \dfrac{n_{(HCl)}}{v_{(HCl)}}$

   $c_{(HCl)} = \dfrac{0.01253}{0.100}$

   $c_{(HCl)} = 0.125 \text{ mol L}^{-1}$ (3 s.f.)

This tells us that, in order to completely neutralise 0.5012 g of sodium hydroxide dissolved in 100 mL of water, we would need 100 mL of 0.125 mol L^{-1} hydrochloric acid.

Science as a human endeavour:

A titration is a procedure used to identify the concentration of a solution by reacting it with a solution of known concentration and measuring the volume required for a complete reaction. An acid–base titration involves determining the concentration of a solution (**analyte**) by a neutralisation reaction with a solution of known concentration (**titrant**). The volume of titrant required to neutralise the analyte is recorded. Titration is a form of **volumetric analysis** given that solution volumes are measured. The progress of the reaction is monitored using an acid–base indicator.

The **equivalence point** of an acid–base reaction is the point at which the concentration of hydronium (H_3O^+) ions is found to be equivalent to the concentration of hydroxide (OH^-) ions. A permanent colour change occurs at the endpoint of the reaction. The **endpoint** is an estimation of the equivalence point and therefore should be close to this value. When completing a titration once the equivalence point is reached, the addition of one more drop gives the approximate endpoint, which is shown by the **first permanent colour change**. The method for completing a titration is explained below.

Consider the following problem: a student is given a 150 mL volumetric flask containing a 0.0500 mol L^{-1} solution of sodium carbonate (Na_2CO_3) and is asked to determine the molar concentration of a hydrochloric acid (HCl) solution using volumetric analysis.

Please refer to Topic 7.6 for the explanation of all glassware used.

1. Obtain a solution of known concentration (the titrant).

   This is done one of two ways. The first is that, as the chemist, you will make up a standard solution using the method given in Topic 4.3.2 and calculate the concentration. The second is that you are given a primary standard; this is the case in this example.

   $c_{(Na_2CO_3)} = 0.0500 \text{ mol L}^{-1}$

2. Rinse all glassware following the guidelines in Topic 7.6.

   In this example, the sodium carbonate is placed into a 50.0 mL burette and a 25.0 mL aliquot of hydrochloric acid is transferred into a conical flask using a volumetric pipette.

3. Take the initial reading for the volume of the burette and record this in a table similar to the one below.

   It is important to note that the resolution of the burette is 0.10 mL.

   Chemists may choose to approximate the volume to the nearest 0.05 mL to improve the precision of the titre values.

	Rough	**1**	**2**	**3**
Initial volume (mL)	0.10			
Final volume (mL)				
Titre (mL)				

4. Transfer 2–3 drops of indicator solution to the hydrochloric acid solution in the conical flask.

   In this example, phenolphthalein will be used. The endpoint will be indicated from a colour change from colourless to a permanent pink tinge.

5. While swirling the conical flask with one hand, use your other hand to turn on the tap of the burette so that the solution is coming out drop by drop. When a colour change starts to become apparent, slow the time between drops down so that the endpoint can be reached upon the first permanent colour change.

6. Measure and record the volume of titrant transferred from the burette to the analyte in the conical flask.

	Rough	**1**	**2**	**3**
Initial volume (mL)	0.10			
Final volume (mL)	24.75			
Titre (mL)	24.65			

7. Repeat this process until three concordant titres are reached.

   The resolution of the instrument (burette) used to transfer the titrant to the analyte is 0.10 mL. The volume markings on the burette increase by 0.10 mL (i.e. 1.10, 1.20, 1.30). Values below 0.10 mL cannot be resolved with confidence. Therefore, 0.10 mL is the minimum requirement for concordant titres.

	Rough	**1**	**2**	**3**
Initial volume (mL)	0.10	2.05	0.00	1.55
Final volume (mL)	24.75	26.55	24.55	26.00
Titre (mL)	24.65 **X**	24.50	24.55	24.45

   The last three titres are concordant (these are within 0.10 mL), whereas the first rough one was not within 0.10 mL of the others.

   An X can be placed next to the titre value that is not going to be used in any subsequent calculations.

8. Deduce the average titre value.

   $$\text{Average titre} = \frac{\text{titre 1} + \text{titre 2} + \text{titre 3}}{3}$$

   $$\text{Average titre} = \frac{24.50 + 24.55 + 24.45}{3}$$

   Average titre = 24.50 mL

   This is the volume of sodium carbonate that will be used in subsequent calculations.

9. Write out the balanced equation for the reaction.

   $$Na_2CO_{3(aq)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

10. Deduce the number of moles of known concentration used in the titration.

   $$c_{(Na_2CO_3)} = \frac{n_{(Na_2CO_3)}}{V_{(Na_2CO_3)}}$$

   $$\therefore\ n_{(Na_2CO_3)} = c_{(Na_2CO_3)} \times V_{(Na_2CO_3)}$$

   $$n_{(Na_2CO_3)} = 0.0500 \times 0.02450$$

   $$n_{(Na_2CO_3)} = 0.001225 \text{ mol}$$

11. Using the mole ratio deduce the number of moles of the unknown solution.

   From the reaction equation we can see that the mole ratio is:

   $$\frac{Na_2CO_3}{1} : \frac{HCl}{2}$$

   The stoichiometric ratio is 1:2, which shows that one mole of sodium carbonate neutralises two moles of hydrochloric acid.

   $$n_{(HCl)} = 2 \times n_{(Na_2CO_3)}$$

   $$n_{(HCl)} = 2 \times 0.001225$$

   $$n_{(HCl)} = 0.00245 \text{ mol}$$

12. Calculate the unknown concentration.

   Each titration was completed using a 25.0 mL (0.025 L) aliquot of hydrochloric acid.

   $$V_{(HCl)} = 25.0 \text{ mL}$$

   $$V_{(HCl)} = 0.0250 \text{ L}$$

   $$c_{(HCl)} = \frac{n_{(HCl)}}{V_{(HCl)}}$$

   $$c_{(HCl)} = \frac{0.00245}{0.0250}$$

   $$c_{(HCl)} = 0.0980 \text{ mol L}^{-1} \text{ (3 s.f.)}$$

Therefore, the concentration of the hydrochloric acid solution is 0.0980 mol L^{-1}.

Questions

22. Determine the volume of 0.103 M hydrochloric acid solution required to neutralise a 50.0 mL volume of 0.0612 M sodium hydroxide solution.

(7 marks) **KA4**

23. A 25.0 mL aliquot of sodium carbonate solution neutralises a 24.55 mL volume of 0.1034 mol L^{-1} hydrochloric acid solution. Determine the concentration of the sodium carbonate solution.

(8 marks) **KA4**

24. Determine the mass of sodium oxide required to neutralise a 50.0 mL volume of 2.013 mol L^{-1} sulfuric acid solution.

(7 marks) **KA4**

25. Determine the maximum mass of sodium chloride produced in the reaction between a 30.0 mL volume of 0.0522 mol L^{-1} sodium carbonate solution with hydrochloric acid solution.

    (a) Write a balanced equation representing the reaction taking place.

    (2 marks) **KA4**

    (b) Calculate the mass of salt produced from the reaction.

    (6 marks) **KA4**

    (c) Explain the process that could be undertaken to separate the salt produced from the solution.

    (2 marks) **KA1**

    (d) Calculate the mass of carbon dioxide gas that would be produced from this reaction.

    (3 marks) **KA4**

5.3: The pH scale

The pH scale

Science understanding

The pH scale is a logarithmic scale that describes the concentration of hydrogen ions in aqueous solutions.

Solutions with pH < 7 are acidic, solutions with pH > 7 are alkaline, and solutions with pH = 7 are neutral.

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The pH scale is used to determine whether a solution is acidic, basic or neutral. The meaning of the letter p in pH is disputed, with scientists disagreeing whether it stands for ‘potential’ or ‘power’. The capital letter H is the elemental symbol for hydrogen.

Science as a human endeavour:

The development of the pH scale is attributed to a number of scientists working independently at the conclusion of the 19th century.

Wilhelm Ostwald discovered a number of dissociation constants for weak acids and used these values to represent the measure of an acid’s strength. Ostwald’s method for measuring the dissociation constants was to dissolve an acid in water and to measure the resulting electrical conductivity of the solution.

Compounds that dissolve in solution and release ions are called **electrolytes**; a compound that completely dissociates is a strong electrolyte and a compound which partially dissociates is a weak electrolyte. Pure water was known to be a non-electrolyte. The conductivity of a solution can be measured with a very high level of precision. Measurements of the electrical conductivity of pure water reveal a very weak electrical conductivity, which shows that even pure water contains a measurable concentration of dissociated ions. Scientists discovered that one in one hundred trillion (10^{14}) water molecules (H_2O) would ionise and form hydronium (H_3O^+) and hydroxide (OH^-) ions. The presence of dissociated ions is explained by the reaction of two molecules of water in solution.

$$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

The ion product (K_w) of pure water is the product of the concentration of hydronium and hydroxide ions in a volume of water. Scientists determined that the ion product of pure water is 1×10^{-14} (at 25°C).

$$K_w = [OH^-].[H_3O^+]$$

$$1 \times 10^{-14} = [OH^-].[H_3O^+]$$

Square brackets [ ] denote concentration in mol L^{-1} and thus $[OH^-]$ represents the hydroxide ion concentration.

The hydrogen ion concentration in pure water was found to be 1×10^{-7} mol L^{-1} and so in 1904 Hans Friedenthal pointed out that the acidity of a solution was related to the solution's hydrogen ion concentration.

$$\frac{1 \times 10^{-14}}{[OH^-]} = 1 \times 10^{-7} \text{ mol L}^{-1}$$

A neutral solution is found to have equal concentrations of both hydronium and hydroxide ions.

$$1 \times 10^{-14} = [OH^-].[H_3O^+]$$

$$\frac{1 \times 10^{-14}}{[H_3O^+]} = [OH^-]$$

$$\frac{1 \times 10^{-14}}{1 \times 10^{-7}} = 1 \times 10^{-7}$$

Science as a human endeavour:

In 1909, Danish biochemist Søren Sørensen was working with amino acids at the Carlsberg Laboratories in Copenhagen. Here, he developed a method for determining the electrical conductivity of amino acids in solution. Sørensen developed a scale based on the ion product of water to determine the degree of ionisation of an amino acid in solution. Sørensen's scale is logarithmic. The pH is calculated by finding the base-10 logarithm of the hydronium ion concentration.

$$\mathrm{pH} = \log [H_3O^+]$$

$$\text{e.g. } \mathrm{pH} = \log(10^{-7})$$

$$\mathrm{pH} = -7$$

Sørensen chose to include a minus sign in his equation such that the solution of the expression is a positive value.

If the concentration of hydronium ions in solution is very low (1×10^{-14}) then the pH is equal to 14. Sørensen chose 14 as the upper limit for his pH scale, as this represents the lowest possible concentration of hydronium ions in solution. The lower limit for the pH scale is set at 0.

Sulfuric acid is a strong acid which has a molar concentration of 18 M and a pH of 0.50. Conversely, sodium hydroxide is a strong base which has a pH between 13.5 and 14.

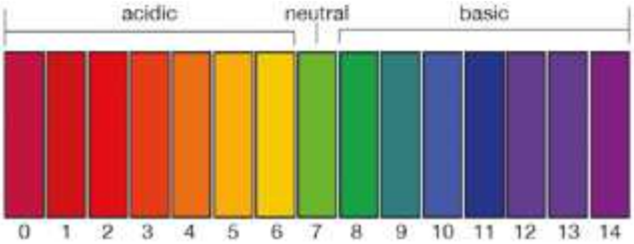

Figure 5.08: Representation of the pH scale when universal indicator is added to the associated solutions.

Questions

26. Consider the following pH values and classify them as relating to an acidic, alkaline or neutral solution.

(a) 5.6. ....

(b) 9.2. ....

(c) 7.. ....

(3 marks) **KA1**

27. Calculate the concentration of hydronium ions in a solution where the concentration of hydroxide ions is 0.015 mol L^{-1}.

(2 marks) **KA4**

28. Calculate the concentration of hydroxide ions in a solution containing 0.123 mol L^{-1} hydronium ions.

(2 marks) **KA4**

pH calculations

Calculations can be completed using the formula that Sørenson developed to determine whether a solution is acidic, basic or neutral based on the concentration of hydronium (H_3O^+) ions.

$$pH = -\log[H_3O^+]$$

Where $[H_3O^+]$ is equal to the concentration of H_3O^+ ions in moles per litre (mol L^{-1}) or (M).

Examples

Consider a 0.05 M sample of hydrochloric acid and calculate the pH of the solution.

$$pH = -\log[H_3O^+]$$

$$pH = -\log[0.05]$$

$$pH = 1.3$$

The pH of a 0.05 M sample of hydrochloric acid is equal to 1.3.

Consider a 0.10 M sample of sodium hydroxide and calculate the pH of the solution.

$$NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

As can be seen from the above equation, sodium hydroxide ionises to produce hydroxide ions so the concentration of hydroxide ions is 0.10 M. We can rearrange the ionic product expression to find the concentration of hydronium ions in this sample.

$$1 \times 10^{-14} = [OH^-][H_3O^+]$$

$$\frac{1 \times 10^{-14}}{[H_3O^+]} = [H_3O^+]$$

$$\frac{1 \times 10^{-14}}{0.1} = [H_3O^+]$$

$$1 \times 10^{-13} = [H_3O^+]$$

Now insert this concentration into the pH equation.

$$pH = -\log[H_3O^+]$$

$$pH = -\log[1 \times 10^{-13}]$$

$$pH = 13$$

The pH of a 0.10 M sample of sodium hydroxide is equal to 13.

To determine the concentration of hydronium ions, the expression can be rearranged as follows.

$$pH = -\log[H_3O^+]$$

$$-pH = \log[H_3O^+]$$

$$10^{-pH} = 10^{\log[H_3O^+]} \quad (10^{\log a} = a)$$

$$10^{\log a} = a$$

Example

Consider a 0.100 M sample of ethanoic acid. According to the pH calculation, this would indicate that ethanoic acid has a pH of 1. However, the pH of the sample is measured as 2.4 using a pH probe. Calculate the concentration of hydronium ions.

$$[H_3O^+] = 10^{-pH}$$

$$[H_3O^+] = 10^{-pH}$$

$$[H_3O^+] = 10^{-2.4}$$

$$[H_3O^+] = 0.0040 \text{ mol L}^{-1}$$

Ethanoic acid is a weak acid, which indicates that it only partially ionises into acetate ions and hydronium ions in solution. In this example, we can observe that in a 0.100 M solution the concentration of hydronium ions is only 0.003981 M. This proves that ethanoic acid only partially ionises. In fact, only 3.9% of ethanoic acid molecules are converted into hydronium ions, and therefore that ethanoic acid is a weak acid (*refer to Figure 5.1.3)*.

Questions

29. Calculate the pH of a 0.0023 mol L^{-1} solution of nitric acid.

(2 marks) **KA4**

30. Calculate the concentration of a sample of hydrochloric acid that has a pH of 2.1.

(2 marks) **KA4**

31. Calculate the pH of a 10^{-4} mol L^{-1} sample of potassium hydroxide.

(4 marks) **KA4**

32. Calculate the concentration of hydroxide ions in a solution with a pH of 8.3.

(4 marks) **KA4**

Acid rain

Science understanding

CO_2 dissolves in rainwater to form carbonic acid, which is a weak acid, giving rainwater a pH of about 5.6.

Oxides of sulfur and nitrogen in the atmosphere can produce rain with a pH below 5.6.

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Carbon dioxide is found in the atmosphere as a result of natural and human processes. Naturally, the ocean releases carbon dioxide into the atmosphere as a result of large amounts of carbon dioxide being dissolved in the ocean. The decomposition of organisms in addition to plant and animal respiration are likewise large producers of carbon dioxide. The respiration process involves glucose ($C_6H_{12}O_6$) and oxygen (O_2) reacting within organisms to produce water (H_2O) and carbon dioxide (CO_2).

$$C_6H_{12}O_{6(s)} + 6O_{(aq)} \rightarrow 6H_2O_{(l)} + 6CO_{2(g)}$$

Carbon dioxide is also produced naturally as a result of volcanoes erupting, but volcanic eruptions produce far less carbon dioxide than the aforementioned processes.

Rainwater is naturally acidic because carbon dioxide (CO_2) in the atmosphere dissolves in the water droplets (H_2O) to produce a weak acid; carbonic acid (H_2CO_3).

$$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$$

Remember that, as it is a weak acid, carbonic acid will only partially ionise in water to produce hydronium ions. This therefore lowers the pH of rainwater. It is for this reason that rainwater naturally has a pH of about 5.6.

$$H_2CO_{3(aq)} + 2H_2O_{(l)} \rightleftharpoons CO^{2-}_{3(aq)} + 2H_3O^+_{(aq)}$$

Acid rain is defined as rainwater with a pH less than 5.6 and occurs when the hydronium ion concentration in water droplets is increased.

The production of acid rain occurs when other stronger acids form as a result of oxides of sulfur and nitrogen dissolving in water droplets in the atmosphere.

Oxides of sulfur

Sulfur is a contaminant found in many fossil fuel sources (oil, petrol) and in carbon used in coal-fired power stations. Most of these sources are used to produce energy and this energy is achieved through burning, with the substances reacting with oxygen gas. As a result of burning these fossil fuels, the contaminant sulfur (S) also reacts with oxygen (O_2) to produce sulfur dioxide (SO_2), an acidic oxide.

$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

Sulfur is found as an impurity in metallic ores and when these ores are roasted sulfur dioxide is released. An example of this is during the production of zinc; zinc sulfide (ZnS) is roasted in oxygen gas (O_2) to remove the sulfur. This process produces sulfur dioxide (SO_2) and zinc oxide (ZnO).

$$2ZnS_{(s)} + 3O_{2(g)} \rightarrow 2ZnO_{(s)} + 2SO_{2(g)}$$

Volcanoes are a natural contributor of sulfur dioxide to the atmosphere. The sulfur dioxide produced by human activity is also released into the atmosphere. Once in the atmosphere, sulfur dioxide (SO_2) dissolves in water droplets (H_2O) and forms sulfurous acid (H_2SO_3), a weak acid.

$$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

Some sulfur dioxide (SO_2) reacts further with oxygen (O_2) in the atmosphere to form sulfur trioxide (SO_3).

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

Sulfur trioxide dissolves in rainwater and forms sulfuric acid, a strong acid. This then precipitates as acid rain.

$$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$$

Oxides of nitrogen

The oxides of nitrogen also contribute to the formation of acid rain. Nitrogen (N_2) can react with oxygen (O_2) to form the colourless gas nitric oxide (NO). This reaction is endothermic and has a high activation energy. Nitric oxide is produced naturally as a result of lightning discharges and bushfires. Within the internal combustion engines of cars and jet engines, and during industrial processes, sufficient energy is released to initiate the formation of nitric oxide. All of these processes occur at the high temperatures (≈1000°C) required to break the strong nitrogen–nitrogen triple-covalent bond present in nitrogen gas (N_2).

$$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$$

The nitric oxide (NO) emitted from these processes then reacts further with oxygen (O_2) in the air to form the brown gas nitrogen dioxide (NO_2).

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$$

In the atmosphere, nitrogen dioxide (NO_2) reacts with water (H_2O) to form a mixture of nitric (HNO_3) and nitrous acid (HNO_2).

$$2NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)} + HNO_{2(aq)}$$

Nitrous acid is a weak acid and therefore does not ionise completely in water. Nitric acid is a strong acid and completely ionises in water to form nitrate and hydronium ions. This reaction can be written as the equation below to better represent what is occurring in the atmosphere.

$$2NO_{2(g)} + 2H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO^-_{3(aq)} + HNO_{2(aq)}$$

When sulfuric and nitric acid ionise in water droplets in the atmosphere, they produce hydronium ions. This results in the pH of rain dropping below 5.6 and so it becomes acidic rain.

$$HNO_{3(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + NO^-_{3(aq)}$$

and

$$H_2SO_{4(aq)} + 2H_2O_{(l)} \rightarrow 2H_3O^+_{(aq)} + SO^{2-}_{4(aq)}$$

Science as a human endeavour

The lower pH of acid rain can have many detrimental effects on the natural and built environment of the world.

As a result of the lower pH of rainwater, plant leaves can lose their waxy coating, causing the leaves to become discoloured and susceptible to bacteria attacks. Acid rain also causes the essential nutrients from the root systems of plants to be **leached** away. Clay particles in soil are negatively charged and attract metal cations to their surface. Many of these metal cations such as magnesium (Mg^{2+}) and potassium (K^+) ions are useful for plants. However, there are also toxic metal cations that are attracted to the clay. Leaching occurs when acid rain falls; the hydrogen ions (H^+) from the rainwater are able to exchange the metal cations (K^+) from the surface of the clay.

$$K^+_{(clay)} + H^+_{(aq)} \rightleftharpoons K^+_{(aq)} + H^+_{(clay)}$$

This results in nutrients being mobilised and washed away so that plant root systems can no longer use them.

Aluminium ions are an example of the types of toxic cations that are attracted to clay. When acidic rain falls, the hydrogen ions (H^+) displace aluminium ions (Al^{3+}) from the surface of the clay.

$$Al^{3+}_{(clay)} + 3H^+_{(aq)} \rightleftharpoons Al^{3+}_{(aq)} + 3H^+_{(clay)}$$

When these cations are displaced from the surface of the clay, they reduce a plant's ability to grow and fight disease. Once displaced, aluminium ions are also able to enter waterways where they can damage fish gills. This can reduce a fish's ability to extract oxygen and therefore sometimes results in death.

We have seen before that acids react with calcium carbonate to produce water, a salt and carbon dioxide. Calcium carbonate ($CaCO_3$) is the compound that makes up limestone and marble. Marble and limestone is used in the building industry and has been used to build many famous structures around the world, including the Pyramids of Giza and medieval churches throughout Europe. Calcium carbonate is an insoluble compound, but, when it reacts with strong acids (H^+) in acid rain, more soluble calcium salts (Ca^{2+}) are formed. This is represented by the following ionic equation.

$$CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

The soluble calcium salts are then washed away with rainwater, leading to the erosion of limestone in the natural environment and in man-made structures around the world.

Acid rain reduces the pH levels of freshwater bodies and, as a result, can be harmful to aquatic life. When the pH of the water reaches 5, fish eggs cannot hatch. Lower pH levels cause some adult fish to die.

Farmers can reduce the effects of acid rain on their fields by adding crushed limestone to the soil. Crushed limestone ($CaCO_3$) helps to neutralise acidic water and therefore increases the pH of the soil. Acid rain can also cause problems for the human respiratory system. These issues are not due to the rainwater, but are because of the fine sulfate and nitrate particles that are produced from interactions in the atmosphere. These particles can be transported over long distances and are then inhaled by humans into their lungs.

In order to reduce the production of acidic rain, strict emission controls have been implemented. These controls focus on reducing the amount of acidic oxides being emitted into the atmosphere. Many governments require scrubbers to be used in industry smoke stacks . These scrubbers contain metal oxides which neutralise the pollutants before they are released into the atmosphere.

Since the 1970s, cars have been required to come equipped with a **catalytic converter**. These convert harmful gases into less harmful gases. Nitric oxide is a gas that contributes to acid rain and is converted within the catalytic converter. A catalytic converter utilises a platinum catalyst (Pt), which catalyses (quickly causes) the reaction between harmful gases. Nitric oxide (NO) will react with carbon monoxide (CO) within the catalytic converter to form carbon dioxide (CO_2) and nitrogen gas (N_2).

$$2CO_{(g)} + 2NO_{(g)} \rightarrow 2CO_{2(g)} + N_{2(g)}$$

Nitric oxide (NO) also reacts with hydrogen gas (H_2) within these converters, and forms water (H_2O) and nitrogen gas (N_2).

$$2H_{2(g)} + 2NO_{(g)} \rightarrow 2H_2O_{(g)} + N_{2(g)}$$

Questions

33. Acid rain can have detrimental effects on humans and the environment.

(a) Oxides of sulfur contribute to the formation of acid rain.

(i) Describe how oxides of sulfur are formed from anthropogenic (human) activities.

(2 marks) **KA2**

(ii) Draw the structural formula for sulfur trioxide and state its shape.

(3 marks) **KA4**

(b) Oxides of nitrogen also contribute to the formation of acid rain.

(i) Define acid rain.

(2 marks) **KA1**

(ii) Write balanced equations representing the formation of acidic rain from nitric oxide.

(1)

(2) *(4 marks)* **KA4**

(c) Explain how acid rain can have detrimental effects on humans and the environment.

(6 marks) **KA3**

Summary Test 5: Acids and bases

1. When soft drink bottles are opened, they produce a fizzing sound caused by the dissolved carbonic acid.

(a) Use a balanced chemical equation to explain where this fizz comes from.

(3 marks) **KA1**

(b) Carbonic acid is considered a weak acid. Define the term 'weak acid'.

(2 marks) **KA1**

(c) State, giving a reason, why a 0.1 mol L^{-1} hydrochloric acid solution has a pH lower than a 0.1 mol L^{-1} carbonic acid solution.

(2 marks) **KA2**

(d) State and explain which solution would have a greater electrical conductivity.

(2 marks) **KA1**

(e) Describe and explain the reaction that takes place when a lit splint is placed over the top of the bottle when it was opened.

(2 marks) **KA2**

(f) Write equations to show the two-stage ionisation of carbonic acid in water.

(4 marks) **KA4**

(g) A student took a sample of the soft drink and placed it onto litmus blue paper. State what colour this paper would be as a result.

(1 mark) **KA1**

(h) Before opening up the bottle, the student set up an apparatus that allowed them to capture all carbon dioxide released from the bottle. The bottle was found to release 2.401 g of carbon dioxide. Calculate the mass of carbonic acid solution required to make this much carbon dioxide.

(5 marks) **KA4**

2. Cloudy ammonia is a common household cleaning product. One company claims that the concentration of ammonia in their product is 20 g L^{-1}.

   (a) Convert the claimed mass concentration of ammonia into a molar concentration (mol L^{-1}).

(2 marks) **KA4**

A student completed a titration to test the claimed concentration of ammonia in the product. The product was too concentrated for what the student was able to complete in the laboratory, so the product was first diluted. The student took a 25.0 mL aliquot of the cloudy ammonia and placed it into a 250 mL volumetric flask. The flask was then filled to the calibration mark with deionised water, stoppered and inverted a few times.

(b) Calculate the theoretical molar concentration of the diluted sample of ammonia.

(2 marks) **KA4**

(c) Explain why the volumetric flask was inverted.

(1 marks) **IAE4**

The sample of ammonia was then used in a titration with a 0.0959 mol L^{-1} standard solution of hydrochloric acid. A 25.0 mL aliquot of the ammonia sample was placed into a conical flask along with 3 drops of methyl orange indicator. The hydrochloric acid was placed into the burette.

(d) Write the balanced chemical equation for the reaction that took place in the conical flask.

(2 marks) **KA4**

The titration was repeated until 3 concordant titres were achieved. The results can be seen in Table A.

	Rough	**1**	**2**	**3**
Initial volume (mL)	0.05	1.00	0.10	3.45
Final volume (mL)	23.70	24.80	23.70	27.10
Titre (mL)	23.65	23.80	23.60	23.65

Table A Titration results

(e) Explain the meaning of the term 'concordant titres'.

(1 mark) **KA1**

(f) Calculate the average titre for this practical.

(1 mark) **KA4**

(g) Calculate the molar concentration, mol L^{-1}, of ammonia in the diluted sample.

(5 marks) **KA4**

(h) Calculate the molar concentration of ammonia in the original product.

(2 marks) **KA4**

(i) Comment on whether the claimed concentration of the ammonia in the commercial product was correct.

(1 mark) **IAE3**

3. Acid rain dissolves carbonate structures.

(a) Define the term 'acid rain'.

(2 marks) **KA1**

(b) Write a fully balanced ionic equation to show the reaction of acid rain with calcium carbonate.

(2 marks) **KA4**

(c) Acid rain dissolves minerals containing calcium, which increases the concentration of dissolved calcium ions which enter the home through the water supply. State one disadvantage of the increase in calcium ions in water.

(1 mark) **KA2**

(d) State one other detrimental effect that acid rain can have on the environment.

(1 mark) **KA1**

4. Oxides of sulfur are found to contribute to acid rain.

   (a) Identify one natural source of oxides of sulfur.

   (1 mark) **KA1**

   (b) Use a balanced equation to show the emission of sulfur dioxide from the roasting of metallic sulfides.

   (2 marks) **KA4**

   (c) Explain with the aid of equations how oxides of sulfur contribute to the formation of acid rain.

   (6 marks) **KA3**

   (d) Draw the structural formula for sulfuric acid.

   (2 marks) **KA4**

   (e) Explain one way in which humans could help to reduce the formation of acid rain.

   (2 marks) **KA3**

Topic 6: Redox Reactions

6.1: Concepts of oxidation and reduction

Science understanding

A range of reactions, including reactions of metals, combustion, and electrochemical processes, can be considered as redox reactions.

The majority of known chemical reactions are reduction-oxidation reactions, or **redox reactions**, and they are occurring all around us, and within us, at all times. Commonly known examples include rusting, burning, smelting, bleaching, the discharge and recharge of batteries, and many reactions that are important in biology (e.g., photosynthesis, respiration).

Figure 6.1: Rusting, burning, smelting, bleaching and photosynthesis are common examples of reduction-oxidation reactions.

Any reaction which involves the **transfer of electrons** between atoms or molecules is a redox reaction.

Example

When octane (a component of petrol) is burnt in air, each carbon atom that was part of the octane molecule combines with oxygen to become part of a carbon dioxide molecule. Likewise, each hydrogen atom combines with oxygen to become part of a water molecule.

$$2\ C_8H_{18(l)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(g)}$$

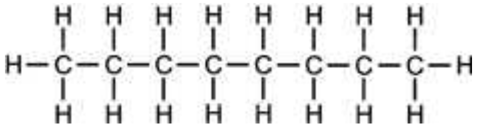

Figure 6.2: Chemical structure of octane

After the reaction, every atom from the octane is bonded with oxygen atoms. The octane has been **oxidised**. Correspondingly, the oxygen has been **reduced**.

Example

Chemical batteries which provide storage and portability for electrical energy are also powered by redox reactions. The specific chemistry depends on the application: portable electronics require lightweight batteries with high energy density, while electric vehicle manufacturers prioritise batteries with long lifespans and excellent safety characteristics. In all cases, the electron transfer of some redox reaction is channelled through a conductive wire so that the energy of the chemical reaction can be extracted as electrical energy.

Figure 6.3: Our understanding of redox chemistry allows us to create portable sources of electricity in a variety of situations.

It is worth emphasising that redox reactions are many and varied. In fact, it would be quicker to identify the reactions which are **not** redox reactions: acid-base reactions and some organic reactions.

Science understanding

Oxidation and reduction can be defined in terms of combination with oxygen, transfer of electrons, or change in oxidation number.

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As chemists' understanding of reactions and atomic structure has become more sophisticated over time, oxidation and reduction have been defined in several different ways.

1a. Historically, oxidation has been defined as oxygen being combined with an element or compound.

Before the discovery of electrons, chemists had already noticed that many substances could combine with oxygen, and they called these 'oxidation reactions'. One such reaction is the burning of iron in air:

$$4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$$

Oxygen has a strong affinity for electrons (i.e., it is highly electronegative), so oxidation reactions almost always result in electrons being transferred from the oxidised substance to the oxygen itself.

1b. Historically, reduction has been defined as oxygen being removed from an element or compound.

Early chemists also knew that many oxidised substances could be broken down into their original components. At high temperatures, mercury(II) oxide decomposes to give elemental mercury and oxygen:

$$2\ HgO \rightarrow 2\ Hg + O_2$$

The products of this reaction are simpler than the reactant. Additionally, the mass of the mercury is less than the mass of the mercury(II) oxide because the oxygen atom is no longer included. For both of these reasons, this process was called 'reduction'.

The modern definitions of oxidation and reduction are more general versions of these older definitions. The advantage of the modern definition is that it allows us to understand reactions that involve the transfer of electrons, but do not involve oxygen.

Science as a human endeavour:

Phlogiston

The slow maturation of understanding of redox reactions is a good example of how scientific hypotheses are developed and discarded. Chemists in the 18th century generally subscribed to the *phlogiston theory*, which posited that combustion involved the release of an invisible elemental substance, phlogiston. Substances that were completely burned were believed to have run out of phlogiston, and the limited capacity of air to support combustion in a closed space was ascribed to phlogiston saturation.

Problems with the phlogiston theory became apparent, however. Most notably, the theory could not explain why the substance produced when metals underwent combustion generally had a greater mass than the initial metal sample. Then, in the late 18th century, English chemist Joseph Priestly discovered oxygen, and French chemist Antoine-Laurent Lavoisier realised that combustion involves the combination of oxygen from the atmosphere with a combustible substance. The phlogiston theory was largely discarded by 1800.

2. Oxidation is a loss of electrons, and reduction is a gain of electrons.

When atoms or molecules experience a net transfer of electrons as part of a reaction, it is said that the substance that has lost electrons has been **oxidised**, and the substance that has gained electrons has been **reduced**.

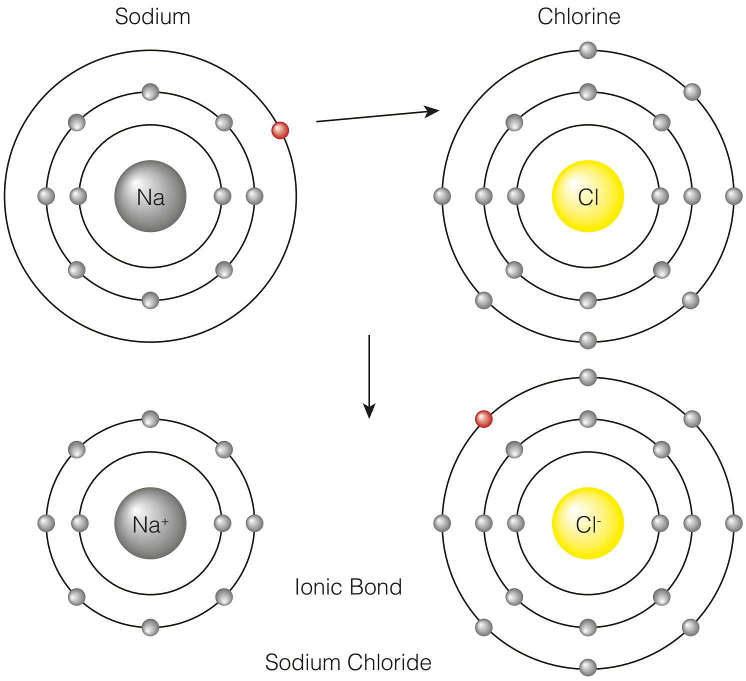

Figure 6.4: Sodium transfers an electron to chlorine; sodium chloride is produced.

For example, sodium metal and chlorine gas react violently to produce sodium chloride (table salt). In this process, the sodium *atom*, which has 11 electrons, donates an electron and is transformed into a sodium *ion*, which has 10 electrons. The sodium has lost an electron (and therefore gained an overall +1 charge).

$$Na \rightarrow Na^+ + e^-$$

The sodium has been oxidised. Chlorine was the **oxidising agent** (or ‘oxidiser’ or ‘oxidant’).

Likewise, the chlorine atom, which has 17 electrons, accepts an electron and is transformed into a chloride ion, which has 18 electrons. The chlorine has gained an electron (and therefore an overall –1 charge).

$$Cl_2 + 2\ e^- \rightarrow 2Cl^-$$

The chlorine has been reduced. Sodium was the **reducing agent** (or ‘reducer’ or ‘reductant’).

Equations such as these, which show only half the redox process, are called **half-equations**. Note that the electron(s) in each equation are the same; they have simply been transferred from sodium to chlorine. It can equally be said that the sodium has reduced the chlorine, or that the chlorine has oxidised the sodium.

A common way of remembering which way the electrons are transferred is to use the mnemonic OIL RIG, which stands for ‘**O**xidation **i**s **L**oss, **R**eduction **i**s **G**ain’.

3. Oxidation is an increase in oxidation number, and reduction is a decrease in oxidation number.

This is the modern definition of reduction and oxidation. Every (non-delocalised) electron can be associated with a specific atom.

If the number of electrons associated with a particular atom is compared to the number present in the elemental form of the substance, the **oxidation number** of that atom can be calculated. If this number changes over the course of a reaction, then the atom has been oxidised or reduced, and the reaction is a redox reaction.

When a sodium atom loses an electron and becomes an ion, its oxidation number increases. Sodium has been oxidised. Note that, by convention, oxidation numbers are written centred above the atom.

$$\overset{\mathbf{0}}{\text{Na}} \rightarrow \overset{\mathbf{+1}}{\text{Na}^+} + e^-$$

Correspondingly, when a chlorine atom gains an electron and becomes an ion, its oxidation number decreases. Chlorine has been reduced.

$$\overset{\mathbf{0}}{\text{Cl}_2} + 2\,e^- \rightarrow 2\,\overset{\mathbf{-1}}{\text{Cl}^-}$$

Oxidation number is **not** simply a tally of how many electrons an atom has; it is calculated by comparison with the unbonded atom, which always has an oxidation number of zero.

Science understanding

Determine the oxidation states of atoms in elements and monatomic ions, and in compounds and polyatomic ions.

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To know whether electrons have been transferred, there must be a method to determine which electrons are associated with which atoms. This is uncomplicated for lone atoms and ionically bonded substances, as they do not share electrons. However, covalently bonded substances seem to pose a problem: covalent bonds consist of shared pairs of electrons, so it can be complicated to determine which atom those electrons ‘belong’ to.

Electrons are equally shared in covalent bonds between identical atoms, so one electron from each bond is assigned to each atom that is bonded.

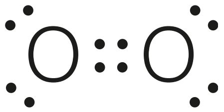

Figure 6.5: Lewis-dot diagram of O_2

The four electrons making up the double-bond are split evenly because the bond is non-polar; two bonding electrons are assigned to each oxygen. This represents no net loss or gain of electrons, so the oxidation number of each oxygen atom is 0.

For bonds between different atoms, all the electrons in the bond are assigned to the more electronegative atom. For example, oxygen is more electronegative than carbon, so in carbon dioxide all the electrons are assigned to the oxygen atoms.

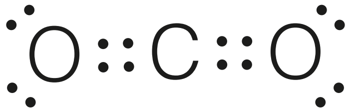

Figure 6.6: Lewis-dot diagram of CO_2

Each oxygen is assigned all four of the bonding electrons. This represents an overall gain of two electrons, so the oxidation number of each oxygen atom is –2. Carbon is not assigned any of the bonding electrons. This represents a loss of four electrons, so the oxidation number of the carbon atom is +4.

Careful! The concept of “assigning” electrons in polar covalent bonds to particular atoms is an abstraction used to keep track of electrons and does **not** represent actual electron exchange (as occurs during ionic bonding). The electrons continue to be unequally shared as part of the covalent bond.

Using these principles, chemists have developed a set of rules for determining oxidation numbers. Rules 1 and 2 are sufficient to determine the oxidation numbers of all atoms in neutral or ionic form. Rules 3 and 4 are necessary to determine oxidation numbers for atoms that are covalently bonded.

1. The oxidation number of an atom in its elemental state is zero.

All pure elements have a formal oxidation number of zero, including both monatomic and diatomic elements.

$$\overset{0}{\mathrm{Na}}\quad \overset{0}{\mathrm{He}}\quad \overset{0}{\mathrm{P}}\quad \overset{0}{\mathrm{H}}{}_2\quad \overset{0}{\mathrm{O}}{}_2\quad \overset{0}{\mathrm{Cl}}{}_2$$

2. The oxidation number of a monatomic ion is equal to the charge of that ion.

$$\overset{+1}{\mathrm{Na}^+}\quad \overset{+2}{\mathrm{Sr}^{2+}}\quad \overset{+3}{\mathrm{Al}^{3+}}\quad \overset{-1}{\mathrm{Cl}^-}\quad \overset{-2}{\mathrm{O}^{2-}}\quad \overset{-3}{\mathrm{Cl}^{3-}}$$

The charge on a monatomic ion can be predicted using the periodic table; this is explained further in Chapter 2. Determining oxidation numbers is therefore straightforward for simple ionic compounds.

$$\overset{+1}{\mathrm{Na}}\overset{-1}{\mathrm{Cl}}\qquad \overset{+1}{\mathrm{K}}{}_2\overset{-2}{\mathrm{O}}\qquad \overset{+3}{\mathrm{Al}}{}_2\overset{-2}{\mathrm{O}}{}_3$$

3. The sum of oxidation numbers in a neutral compound is zero, and the sum of oxidation numbers in a polyatomic ion is equal to the charge of that ion.

The third rule relates to substances with covalent bonding. When combined with a set of rules about the typical oxidation numbers of specific atoms, it can be used to determine the oxidation numbers of other atoms in the molecule or ion. The rules about typical oxidation numbers are given below.

Example

Carbon Dioxide

The carbon dioxide molecule consists of two oxygen atoms and one carbon atom. It is a neutral compound, so the sum of the oxidation numbers must be 0 (Rule 3). Oxygen typically has an oxidation number of –2 (Rule 4e). For the sum of oxidation numbers to be zero, carbon must have an oxidation number of +4.

$$\overset{+4}{\mathrm{C}}\overset{-2}{\mathrm{O}}{}_2\qquad \overset{-2}{\mathrm{O}}=\overset{+4}{\mathrm{C}}=\overset{-2}{\mathrm{O}}$$

Figure 6.7: Formula and structure of carbon dioxide showing oxidation numbers

Example

Water

The water molecule consists of two hydrogen atoms and one oxygen atom. It is a neutral compound, so the sum of the oxidation numbers must be 0 (Rule 3). Hydrogen typically has an oxidation number of +1 (Rule 4d). For the sum of oxidation numbers to be zero, oxygen must have an oxidation number of –2 (which also matches with Rule 4e).

$$\overset{+1}{\mathrm{H}}{}_2\overset{-2}{\mathrm{O}}$$

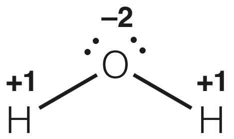

Figure 6.8: Formula and structure of water showing oxidation numbers

Example

Sulfate

The sulfate ion consists of one sulfur and four oxygen atoms. It has an overall charge of 2–, so the sum of oxidation numbers must be –2 (Rule 3). Oxygen typically has an oxidation number of –2 (Rule 4e). For the sum of oxidation numbers to be –2, the sulfur must have an oxidation number of +6.

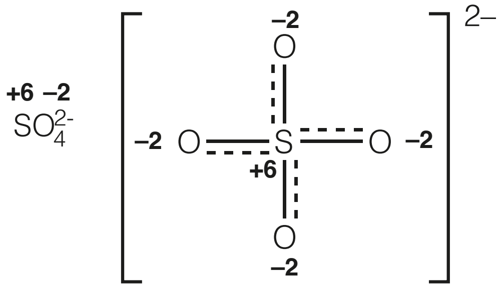
$\overset{+6}{S}\overset{-2}{O}_4^{2-}$

Figure 6.9: Formula and structure of sulfate showing oxidation numbers

4. Some atoms have typical oxidation numbers.

4a. The oxidation number of fluoride is *always* –1.

4b. The oxidation number of ions of metals in Group I is *always* +1.

Example

Lithium Hydride

The lithium hydride molecule consists of one hydrogen and one lithium atom. It is a neutral compound, so the sum of oxidation numbers must be 0 (Rule 3). Lithium always has an oxidation number of +1 (Rule 4b). For the sum of oxidation numbers to be 0, the hydrogen must have an oxidation number of –1. *(Rule 4d says that hydrogen typically has an oxidation number of +1, but Rule 4b takes precedence).*

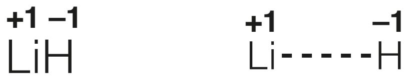
$\overset{+1}{Li}\overset{-1}{H}$

Figure 6.10: Formula and structure of lithium hydride showing oxidation numbers

4c. The oxidation number of ions of metals in Group II is *always* +2.

4d. The oxidation number of hydrogen is *typically* +1.

Exception: when hydrogen is bonded to a Group I or II metal as part of a metal hydride, Rules 4b and 4c take precedence, so hydrogen's oxidation number is –1.

Example

Methane

The methane molecule consists of one carbon and four hydrogen atoms. It is a neutral compound, so the sum of oxidation numbers must be 0 (Rule 3). Hydrogen typically has an oxidation number of +1 (Rule 4d). For the sum of oxidation numbers to be zero, carbon must have an oxidation number of –4.

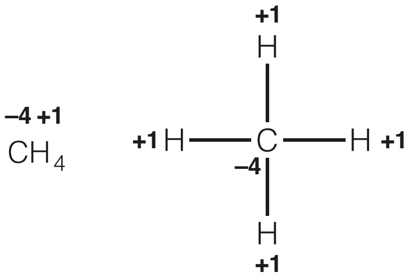
$\overset{-4}{C}\overset{+1}{H}_4$

Figure 6.11: Formula and structure of methane showing oxidation numbers

4e. The oxidation number of oxygen is *typically* –2.

Exceptions: when an atom of oxygen is part of a peroxide group, its oxidation number is –1. When oxygen is bonded to fluorine, its oxidation number is +1.

Example

Hydrogen Peroxide

The hydrogen peroxide molecule consists of two hydrogen and two oxygen atoms. It is a neutral compound, so the sum of oxidation numbers must be 0 (Rule 3). Hydrogen typically has an oxidation number of +1 (Rule 4d). When oxygen is part of a peroxide, it has an oxidation number of –1 (Rule 4e).

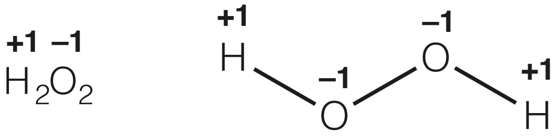

Figure 6.12: Formula and structure of peroxide showing oxidation numbers

Example

Hypochlorous Acid

The hypochlorous acid molecule consists of one hydrogen, one oxygen, and one chlorine atom. It is a neutral compound, so the sum of oxidation numbers must be 0 (Rule 3). Hydrogen typically has an oxidation number of +1 (Rule 4d). Oxygen typically has an oxidation number of –2 (Rule 4e). For the sum of oxidation numbers to be 0, the chlorine must have an oxidation number of +1. *(Rule 4f says that chlorine typically has an oxidation number of –1, but Rules 4d and 4e take precedence).*

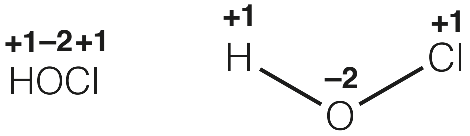

Figure 6.13: Formula and structure of hypochlorous acid showing oxidation numbers

4f. The oxidation number of halogen anions other than fluoride (i.e., chloride, bromide, iodide, astatide, and tennesside) is *typically* –1.

Exception: when a halogen other than fluorine is bonded to fluorine or oxygen, Rules 4a and 4e take precedence.

The complete set of rules is summarised below.

1.	The oxidation number of an atom in its elemental state is zero.
2.	The oxidation number of a monatomic ion is equal to the charge of that ion.
3.	The sum of oxidation numbers in a neutral compound is zero, and the sum of oxidation numbers in a polyatomic ion is equal to the charge of that ion.
4a.	The oxidation number of fluoride is *always* –1.
4b.	The oxidation number of ions of metals in Group I is *always* +1.
4c.	The oxidation number of ions of metals in Group II is *always* +2.
4d.	The oxidation number of hydrogen is *typically* +1.
4e.	The oxidation number of oxygen is *typically* –2.
4f.	The oxidation number of halogen anions other than fluoride (i.e., chloride, bromide, iodide, astatide, and tennesside) is *typically* –1.

Questions

1. Determine the oxidation number of bromine in the following compounds and ions.
   - (a) BrO^- ....
   - (b) BrO_2 ....
   - (c) $NaBrO_4$ ....
   - (d) $HBrO_2$ ....
   - (e) $KBrO_3$ ....
   - (f) Br^- ....

   (6 marks) **KA1**

2. Determine the oxidation number of phosphorus in the following compounds and ions.
   - (a) Ca_3P_2 ....
   - (b) H_3PO_3 ....
   - (c) $H_4P_2O_6$ ....
   - (d) PO_4^{3-} ....
   - (e) P_2H_4 ....
   - (f) H_3PO_2 ....

   (6 marks) **KA1**

3. Determine the oxidation number of each atom in the following compounds and ions.
   - (a) LiCl .... *(2 marks)* **KA1**
   - (b) O_3 .... *(1 mark)* **KA1**
   - (c) NH_3 .... *(2 marks)* **KA1**
   - (d) NaH .... *(2 marks)* **KA1**
   - (e) CH_3OOH (methyl hydroperoxide) .... *(3 marks)* **KA1**
   - (f) OF_2 .... *(2 marks)* **KA1**
   - (g) HCO_3^- .... *(3 marks)* **KA1**
   - (h) H_2SO_4 .... *(3 marks)* **KA1**
   - (i) $Al(OH)_3$ .... *(3 marks)* **KA1**
   - (j) MnO_2 .... *(2 marks)* **KA1**

4. Identify the species that have been oxidised and reduced in the following reaction.

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

   - (a) Oxidised species: .... *(1 mark)* **KA1**
   - (b) Reduced species: .... *(1 mark)* **KA1**

5. Identify the oxidising agent and reducing agent in the following reaction.

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

   - (a) Oxidising agent: .... *(1 mark)* **KA1**
   - (b) Reducing agent: .... *(1 mark)* **KA1**

6. Identify the species that have lost and gained electrons in the following reaction.

$$4Fe(OH)_2 + O_2 + 2H_2O \rightarrow 4Fe(OH)_3$$

   - (a) Species that lost electrons: .... *(1 mark)* **KA1**
   - (b) Species that gained electrons: .... *(1 mark)* **KA1**

7. Complete the following sentences using appropriate chemical terminology.
   - (a) A loss of electrons is called .... *(1 mark)* **KA4**
   - (b) .... reactions involve the transfer of electrons. *(1 mark)* **KA4**
   - (c) The substance that gains electrons is ....; it is called the .... agent. *(2 marks)* **KA4**
   - (d) The oxidation number of the substance that is .... goes down. *(1 mark)* **KA4**

Science understanding

Identify oxidation and reduction in given equations.

Redox reactions involve two complementary processes: oxidation and reduction. Oxidation is a loss of electrons, and the oxidation number of a species that is oxidised will increase. Reduction is a gain of electrons, and the oxidation number of a species that is reduced will decrease. Oxidation numbers can therefore be used to identify whether a particular reaction is a redox reaction, and which species have been oxidised or reduced.

Consider the reaction of potassium oxide and hydrochloric acid, which produces potassium chloride and water. The reaction is shown below, including the oxidation number of each atom in both the reactants and products.

$$\overset{+2}{K}_2\overset{-2}{O} + 2\,\overset{+1}{H}\overset{-1}{Cl} \rightarrow 2\,\overset{+2}{K}\overset{-1}{Cl}_2 + \overset{+1}{H}_2\overset{-2}{O}$$

By comparison of the oxidation number of each type of atom in the equation both before and after the reaction occurs, it is apparent that no atom has changed oxidation number. For example, the hydrogen atoms have an oxidation number of +1 on both the reactant side and the product side. No atoms have changed oxidation number. This means **the reaction above is not a redox reaction**. It is a neutralisation reaction, a type of acid-base reaction.

Now consider the combustion of hydrogen in oxygen, the product of which is water.

$$2\overset{0}{H}_2 + \overset{0}{O}_2 \rightarrow 2\overset{+1}{H}_2\overset{-2}{O}$$

Here, at least one atom *has* changed its oxidation number, so this *is* a redox reaction. The magnitude and direction of the change in oxidation number helps us identify which substances have been oxidised and reduced. Remember: oxidation is an increase is oxidation number, while reduction is a decrease in oxidation number.

The oxidation number of the hydrogen atom has changed from 0 to +1, so it has increased. The hydrogen has been oxidised by the oxygen. Conversely, the oxidation number of the oxygen atom has changed from 0 to –2, so it has decreased. The oxygen has been reduced by the hydrogen.

A substance which causes oxidation is called an **oxidising agent**. In the reaction above, hydrogen has been oxidised by oxygen, so oxygen is the oxidising agent. In the process of causing oxidation, oxygen has been reduced. This is always the case: **the oxidising agent is itself reduced**.

The same is true for **reducing agents**. In the reaction above, oxygen has been reduced by hydrogen, so hydrogen is the reducing agent. In the process of causing reduction, hydrogen has been oxidised. **The reducing agent is itself oxidised**.

Example

Solid iron will react with aqueous copper(II) sulfate to produce metallic copper and aqueous iron(II) sulfate.

$$\overset{0}{Fe} + \overset{+2}{Cu}\overset{+6}{S}\overset{-2}{O}_4 \rightarrow \overset{+2}{Fe}\overset{+6}{S}\overset{-2}{O}_4 + \overset{0}{Cu}$$

The oxidation numbers of both the iron and the copper have changed, so this is a redox reaction. The oxidation number of iron has changed from 0 to +2, an increase, so iron has been oxidised. The oxidation number of copper has changed from +2 to 0, a decrease, so copper has been reduced.

This can be expressed in several ways. The following statements are all equivalent and correct.

- The copper has oxidised the iron.
- Iron has been oxidised by copper.
- The iron has reduced the copper.
- Copper has been reduced by iron.
- Electrons have been transferred from iron to copper.

Science understanding

Write oxidation and reduction half-equations, in neutral and acidic conditions, given reactant and product species.

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Redox reactions involve the transfer of electrons from one substance to another; reduction and oxidation occur simultaneously. Despite this, it is useful to be able to describe the two processes separately. This is done using **redox half-equations**, which show either the oxidation process or the reduction process that occurs in a reaction. Half-equations are easily recognised because they always show free electrons.

To write an oxidation or reduction half-equation, the relevant **conjugate redox pair** must be identified. Consider the reaction of tin and bromine to give tin(II) bromide.

$$\overset{0}{\mathrm{Sn}} + \overset{0}{\mathrm{Br_2}} \rightarrow \overset{+2\;-1}{\mathrm{SnBr_2}}$$

The conjugate redox pair for the *oxidation* half-reaction is the atom or group of atoms that has *increased* its oxidation number, while the conjugate redox pair for the *reduction* half-reaction is the atom of group of atoms that has *decreased* its oxidation number.

conjugate oxidation pair	Sn/Sn^{2+}
conjugate reduction pair	Br_2/Br^-

Redox Half-Equations in Neutral Conditions

Consider the reaction of iron with chlorine, the product of which is iron(III) chloride.

$$2Fe + 3Cl_2 \rightarrow 2FeCl_3$$

Electrons have been transferred from the iron to the chlorine, producing Fe^{3+} and Cl^- ions. To write two half-equations, the oxidising agent and the reducing agent must first be identified using the oxidation numbers of each atom.

$$\overset{0}{2\mathrm{Fe}} + \overset{0}{3\mathrm{Cl_2}} \rightarrow \overset{+3\;-1}{2\mathrm{FeCl_3}}$$

The oxidation number of iron has increased by 3, so iron has been oxidised (it is the reducing agent). To write the first half-equation, begin by writing the conjugate redox pair as reactant and product.

$$Fe \rightarrow Fe^{3+}$$

Then, electrons are added to one side of the equation to ensure the equation is **charge-balanced** (the same amount of charge on both sides). Electrons are negatively charged, so they must be added to the side that has a higher charge: in this case, the product side. Iron has been oxidised, so it has *lost* electrons, which means the electrons must appear on the right.

The complete **oxidation half-equation** is shown below. Electrons have very close to zero mass, so their addition does not disrupt the mass balance of the equation.

$$Fe \rightarrow Fe^{3+} + 3e^-$$

The same process is followed to write the reduction half-equation.

Step 1. Write the conjugate redox pair as reactant and product, ensuring that the equation is **mass-balanced** (the same number of each atom on both sides).

$$Cl_2 \rightarrow 2Cl^-$$

Step 2. Add electrons to balance the charge. In reduction half-equations, electrons appear on the left.

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

Redox Half-Equations in Acidic Conditions

Some aqueous redox reactions will only occur in acidic conditions because protons (H^+) are required to accept surplus oxygen atoms to form water molecules. These half-equations are more complex, but the acidic conditions mean protons (H^+) and water (H_2O) can be added to the reaction equation at will to ensure that the number of hydrogen atoms and oxygen atoms is balanced.

The process for balancing a half-equation in acidic conditions has four key steps:

1. Balance all atoms other than hydrogen and oxygen.
2. Add water molecules to balance oxygen atoms.
3. Add protons (H^+) to balance hydrogen atoms.
4. Add electrons to balance charge.

Example

In acidified solution, iron(II) ions react with permanganate ions (MnO_4^-) to produce iron(III) and manganese(II) ions. The unbalanced equation is shown below, with oxidation numbers included.

$$\overset{+2}{Fe}{}^{2+} + \overset{+7}{Mn}\overset{-2}{O}{}_4^- \rightarrow \overset{+3}{Fe}{}^{3+} + \overset{+2}{Mn}{}^{2+}$$

The iron has lost an electron, and its oxidation number has increased by 1, so the iron has been oxidised. The oxidation half-equation is straightforward to write.

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

The manganese has gained five electrons, and it oxidation number has decreased by 5, so the manganese has been reduced. (*Careful!* It would be easy to become confused by the charges on the ions, MnO_4^- and Mn^{2+}, and think that electrons have been lost. Pay attention to the oxidation numbers of individual atoms).

$$MnO_4^- \rightarrow Mn^{2+}$$

The process given above can be applied to write a balanced reduction half-equation.

1. Balance all atoms other than hydrogen and oxygen.

Other than the oxygen atoms, the equation is already mass-balanced.

2. Add water molecules to balance oxygen atoms.

There are four oxygens on the left, so four water molecules must be added on the right.

$$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$$

3. Add protons to balance hydrogen atoms.

There are now eight hydrogens on the right, in the water molecules, so eight protons must be added on the left.

$$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$$

4. Add electrons to balance charge.

The charge is $-1 + (8 \times +1) = +7$ on the left, and +2 on the right. Five electrons must be added on the left.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

The reduction half-equation is complete. It shows that, if a reducing agent is available to donate five electrons to the manganese atom, eight protons will combine with the four oxygen ions to produce four water molecules. The permanganate ion, stripped of its oxygens but provided with electrons, becomes a manganese(II) ion.

Questions

8. Phosphorus burns in oxygen.

$$P_4 + 5O_2 \rightarrow P_4O_{10}$$

(a) Identify the change in the oxidation number of oxygen.

.. .. *(1 mark)* **KA1**

(b) Hence, state whether oxygen has been oxidised or reduced.

.. .. *(1 mark)* **KA1**

(c) Identify the change in the oxidation number of phosphorus.

.. .. *(1 mark)* **KA1**

(d) Write a half-equation for the oxidation of phosphorus in this reaction.

(2 marks) **KA1**

(e) State the evidence from the equation that phosphorus has been oxidised (other than the change in oxidation number).

.. .. *(1 mark)* **KA2**

9. Plants use photosynthesis to convert sunlight, water, and carbon dioxide into sugar and oxygen.

(a) Identify the oxidation number of each atom in the photosynthesis equation.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

(4 marks) **KA1**

(b) Identify the species that has been oxidised in the reaction.

.. .. *(1 mark)* **KA1**

(c) Write a half-equation for the reduction of CO_2 to $C_6H_{12}O_6$ in acidic conditions.

(2 marks) **KA1**

Science understanding

Combine half-equations to write a chemical equation.

When both half-equations for a redox reaction have been determined, a fully balanced reaction equation can be written. For the reaction of iron and chlorine to produce iron(III) chloride, the following two half-equations were written:

$$\mathrm{Fe} \rightarrow \mathrm{Fe}^{3+} + 3e^-$$

$$\mathrm{Cl}_2 + 2e^- \rightarrow 2\mathrm{Cl}^-$$

The process for writing a full equation involves three steps.

1. Multiply one or both half-equations so that the number of free electrons is equal.

The lowest common denominator of 2 and 3 is 6, so there must be six free electrons in each half-equation before they are combined. This is achieved by multiplying each equation by the number of electrons in the other equation. This could be called "cross-multiplying" the equations.

$$(\mathrm{Fe} \rightarrow \mathrm{Fe}^{3+} + 3e^-) \times 2$$

$$2\mathrm{Fe} \rightarrow 2\mathrm{Fe}^{3+} + 6e^-$$

and

$$(\mathrm{Cl}_2 + 2e^- \rightarrow 2\mathrm{Cl}^-) \times 3$$

$$3\mathrm{Cl}_2 + 6e^- \rightarrow 6\mathrm{Cl}^-$$

2. Combine the two equations.

The two cross-multiplied equations are summed: the reactants from both equations are written on the reactant side, and the products from both equations are written on the product side.

$$\begin{array}{lcl} 2\mathrm{Fe} & \rightarrow & 2\mathrm{Fe}^{3+} + 6e^- \\ 3\mathrm{Cl}_2 + 6e^- & \rightarrow & 6\mathrm{Cl}^- \\ 2\mathrm{Fe} + 3\mathrm{Cl}_2 + 6e^- & \rightarrow & 2\mathrm{Fe}^{3+} + 6e^- + 6\mathrm{Cl}^- \end{array}$$

3. Cancel the electrons.

The final equation should not contain any free electrons. There are an equal number of free electrons on each side; these can be cancelled and removed.

$$2\mathrm{Fe} + 3\mathrm{Cl}_2 \cancel{+ 6e^-} \rightarrow 2\mathrm{Fe}^{3+} \cancel{+ 6e^-} + 6\mathrm{Cl}^-$$

$$2\mathrm{Fe} + 3\mathrm{Cl}_2 \rightarrow 2\mathrm{Fe}^{3+} + 6\mathrm{Cl}^-$$

which can also be written as

$$2\mathrm{Fe} + 3\mathrm{Cl}_2 \rightarrow 2\mathrm{FeCl}_3$$

This is the fully balanced equation that was originally presented on page 252.

The three-step process for combining half-equations also works for reactions in acidic conditions.

1. Multiply one or both half-equations so that the number of free electrons is equal.

$$\mathrm{MnO}_4^- + 8\mathrm{H}^+ + 5e^- \rightarrow \mathrm{Mn}^{2+} + 4\mathrm{H}_2\mathrm{O}$$

$$(\mathrm{Fe}^{2+} \rightarrow \mathrm{Fe}^{3+} + e^-) \times 5$$

$$5\mathrm{Fe}^{2+} \rightarrow 5\mathrm{Fe}^{3+} + 5e^-$$

2. Combine the two equations.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

$$MnO_4^- + 8H^+ + 5e^- + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^-$$

3. Cancel the electrons.

$$MnO_4^- + 8H^+ \cancel{+ 5e^-} + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} \cancel{+ 5e^-}$$

The full balanced equation is:

$$MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$$

Questions

10. Explain the process of reduction in terms of electron transfer.

....

....

.... *(1 mark)* **KA4**

11. Chlorine reacts with sodium bromide to produce bromine and sodium chloride.

$$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$$

(a) Identify the change in the oxidation number of bromine.

.... *(1 mark)* **KA1**

(b) Hence, state whether bromine has been oxidised or reduced.

.... *(1 mark)* **KA1**

(c) Write a half-equation for the oxidation of bromine in this reaction.

(2 marks) **KA1**

(d) Identify the change in the oxidation number of chlorine.

.... *(1 mark)* **KA1**

12. Aluminium (Al) and iodine (I_2) react to form aluminium iodide (AlI_3).

(a) Write two half-equations to show the processes of oxidation and reduction in this reaction.

(4 marks) **KA1**

(b) Combine the half-equations to write the overall equation. Show all working.

(2 marks) **KA1**

(c) Name the reactant that is oxidised in this reaction.

.. . . *(1 mark)* **KA1**

13. Chromate anions ($Cr_2O_7^{2-}$) can be reduced to chromium(III) ions (Cr^{3+}) in acidic conditions.

$$Cr_2O_7^{2-} \rightarrow Cr^{3+}$$

(a) Rewrite the half-equation, balancing the number of chromium atoms.

(1 mark) **KA1**

(b) Rewrite the half-equation, adding water molecules to balance the number of oxygen atoms.

(1 mark) **KA1**

(c) Rewrite the half-equation, adding protons (H^+) to balance the number of hydrogen atoms.

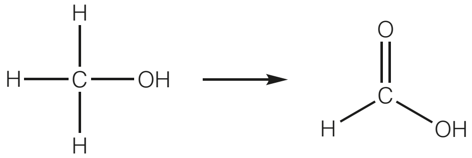
(1 mark) **KA1**

(d) Rewrite the half-equation, adding electrons to balance the charge.

(1 mark) **KA1**

One possible reducing agent for this reaction is methanol (CH_3OH). When acting as a reducing agent in acidic conditions, methanol is itself oxidised to methanoic acid (HCOOH). Methanoic acid has the common name ‘formic acid’.

$$CH_3OH \rightarrow HCOOH$$



The number of carbons in this half-equation is already balanced.

(e) Rewrite the oxidation half-equation, adding water molecules, protons, and electrons to balance the oxygen atoms, hydrogen atoms, and charge.

(1 mark) **KA1**

(f) Multiply the two half-equations so that the number of electrons in each is equal.

(2 marks) **KA1**

(g) Combine the two half-equations, then cancel the electrons and any excess protons or water molecules, to write a full balanced redox equation for the oxidation of methanol by chromate anions.

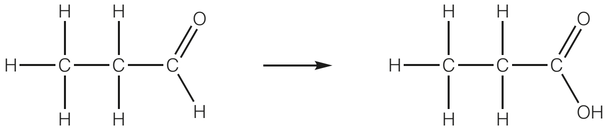
(2 marks) **KA1**

14. Chromium trioxide (CrO_3) is an excellent oxidising agent.

$$CrO_3 \rightarrow Cr^{3+}$$

It can be used to oxidise propanal to propanoic acid in acidic conditions.

$$CH_3CH_2CHO \rightarrow CH_3CH_2COOH$$



(a) Write a balanced half-equation for the reduction of chromium trioxide to chromium(III) ions in acidic conditions.

(2 marks) **KA1**

(b) Write a balanced half-equation for the oxidation of propanal to propanoic acid in acidic conditions.

(2 marks) **KA1**

(c) Combine the half-equations to write a balanced overall equation for the reaction of chromium trioxide and propanal in acidic conditions to give chromium(III) ions and propanoic acid.

Overall equation:

(2 marks) **KA1**

6.2: Metal reactivity

Science understanding

Metals differ in their tendency to lose electrons; more reactive metals lose electrons more easily.

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Different substances have different **reactivities**, which means they have different propensities to react. The reactivity of specific elements depends heavily on the electronic configuration of the atoms and how close they are to attaining a complete outer shell. For example, fluorine is extremely reactive because it requires only one additional electron to complete its octet. Conversely, xenon almost never reacts because it already possesses a complete outer shell, which is a highly energetically stable arrangement (i.e., energy would be required to either add or remove an electron from xenon).

Metals are characterised by their tendency to lose (not gain) electrons to achieve a complete outer shell. Differences in metal reactivity can be understood using electronegativity. All metals would prefer to lose rather than gain electrons, so all metals have relatively low electronegativity. However, less reactive metals have *moderately* low electronegativity, while more reactive metals have *extremely* low electronegativity.

Notice that, for metals, a low electronegativity leads to high reactivity. *The ability of metals to lose electrons, form cations, and participate in ionic bonding is explained further in Topic 2.2.*

Science understanding

A more reactive metal is able to donate electrons to the ion of a less active metal in a displacement reaction.

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A displacement reaction is a reaction in which one element (or polyatomic ion) displaces another in a compound. Most often, this involves a more reactive metal displacing a less reactive metal in an ionic compound.

Example

When a piece of iron metal (Fe) is placed in a solution of copper(II) sulfate ($CuSO_4$), the blue colour of the solution begins to fade and the iron starts to appear copper-coloured. Iron is displacing the copper from the copper(II) sulfate, reducing copper(II) ions (Cu^{2+}), which give the solution its blue colour, into metallic copper (Cu) on the surface of the nail. Simultaneously, iron(II) ions (Fe^{2+}) are released into solution to form aqueous iron(II) sulfate ($FeSO_4$), which is colourless.

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + FeSO_{4(aq)}$$

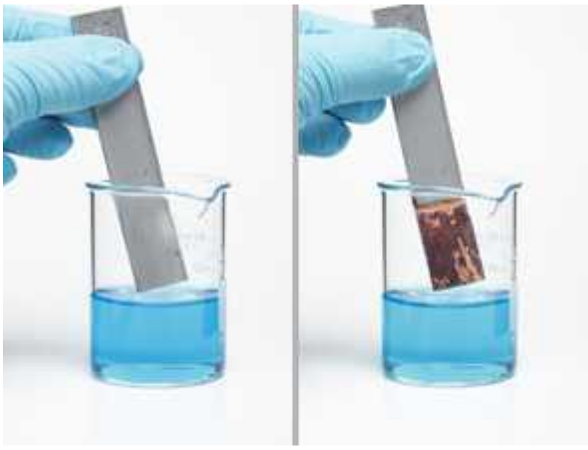

Figure 6.14: Iron reacts with copper sulfate

This occurs because iron is more reactive than copper. Copper has a higher electronegativity than iron, so iron atoms can donate electron to copper atoms. This produces elemental copper and soluble Fe^{2+} ions, which become associated with the sulfate ions.

The sulfate ions do not actually participate in the reaction; they are **spectator ions.** The displacement reaction can also be written as:

$$Fe_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cu_{(s)} + Fe^{2+}_{(aq)}$$

Science understanding

Write equations and half-equations for reactions between a metal and the ion of a less active metal.

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Metal displacement reactions are a type of redox reaction because they involve the transfer of electrons. Half-equations can be written by including only the metal that is being oxidised or reduced, and then balancing the charge by adding electrons to one side of the equation. It is not necessary to know which spectator ions are present to write a pair of redox half-equations for a metal displacement reaction.

Example

Write the oxidation and reduction half-equations for the displacement of copper ions from copper(II) sulfate by iron.

First, identify the oxidation numbers for each atom.

$$\overset{0}{Fe} + \overset{+2}{Cu}\overset{+6}{S}\overset{-2}{O_4} \rightarrow \overset{0}{Cu} + \overset{+2}{Fe}\overset{+6}{S}\overset{-2}{O_4}$$

Iron's oxidation number has increased from 0 to +2. It has been oxidised.

$$Fe \rightarrow Fe^{2+} + 2e^-$$

Copper's oxidation number has decreased from +2 to 0. It has been reduced.

$$Cu^{2+} + 2e^- \rightarrow Cu$$

Example

Write the half-equations for the reduction of lead cations by zinc.

$$\overset{0}{Zn} + \overset{+2}{Pb^{2+}} \rightarrow \overset{+2}{Zn^{2+}} + \overset{0}{Pb}$$

The oxidation number of zinc has increased from 0 to +2. It has been oxidised.

$$Zn \rightarrow Zn^{2+} + 2e^-$$

The oxidation number of lead has decreased from +2 to 0. It has been reduced.

$$Pb^{2+} + 2e^- \rightarrow Pb$$

The above reaction would occur if a sample of zinc was placed in a solution of lead(II) nitrate, for example.

$$Zn_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Pb_{(s)} + Zn(NO_3)_{2(aq)}$$

Questions

15. A reaction occurred when a copper wire (Cu) was placed in silver(I) nitrate solution ($AgNO_3$). The products were copper(II) nitrate and silver (Ag).

    (a) Write the empirical formula for copper(II) nitrate.

    .... *(1 mark)* **KA1**

    (b) Identify the reactant that is oxidised in the reaction.

    .... *(1 mark)* **KA1**

    (c) Write a half-equation to show the oxidation that occurs in this reaction.

    (2 marks) **KA1**

    (d) Write a half-equation to show the reduction that occurs in this reaction.

    (2 marks) **KA1**

(e) Write a net ionic equation for the overall reaction, showing only the species that were involved in the transfer of electrons.

(1 mark) **KA1**

(f) Write a fully balanced chemical equation for the overall reaction.

(1 mark) **KA1**

(g) State what would be observed if gold wire were placed in the solution.

(1 mark) **KA2**

(h) Explain the prediction you made in Question 12 (g).

(1 mark) **KA1**

Science understanding

Differences in metal reactivity can be represented as a metal activity series.

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The metallic elements can be ordered by decreasing reactivity to obtain a **metal activity series**.

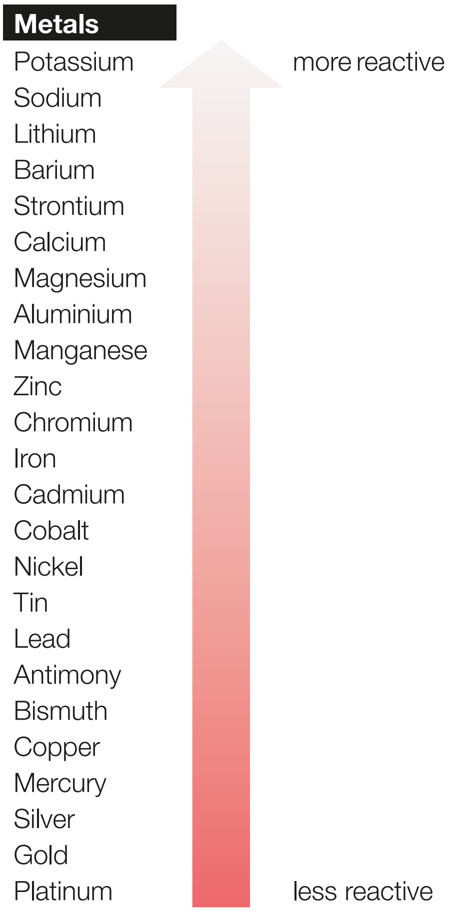

Figure 6.15: Metal activity series

Science understanding

Determine whether a reaction will occur between a metal and a solution containing the ions of another metal, given a metal activity series containing both metals.

Less reactive metals cannot donate electrons to more reactive metals, so displacement reactions only occur when neutral atoms of a more reactive metal can donate electrons to cations of a less reactive metal.

Electrons are only donated *downwards* in the metal activity series, never upwards.

Example

A sample of metallic magnesium placed in a solution of nickel(II) nitrate will react because magnesium is more reactive than nickel.

$$Mg_{(s)} + Ni(NO_3)_{2(aq)} \rightarrow Ni_{(s)} + Mg(NO_3)_{2(aq)}$$

However, a sample of metallic gold placed in copper(II) sulfate will not react because gold is less reactive than copper.

$$Au_{(s)} + CuSO_{4(aq)} \rightarrow \text{no reaction}$$

Questions

16. Using the metal activity series, state whether a reaction would occur between the following substances. *You do not need to write an equation.*

    (a) Lead(II) nitrate solution and solid iron.

    .... *(1 mark)* **KA1**

    (b) Nickel(II) sulfate solution and solid copper.

    .... *(1 mark)* **KA1**

    (c) Copper(II) sulfate solution and solid magnesium.

    .... *(1 mark)* **KA1**

Science understanding

The reactivity of a metal affects its ability to react with other chemicals.

Investigate the reactions of various metals with water and acidic solutions.

Compare the vigour of reactions of different metals with their position on the metal activity series.

The metal activity series is derived by comparing the reactions that each metal will undergo and the rate of these reactions.

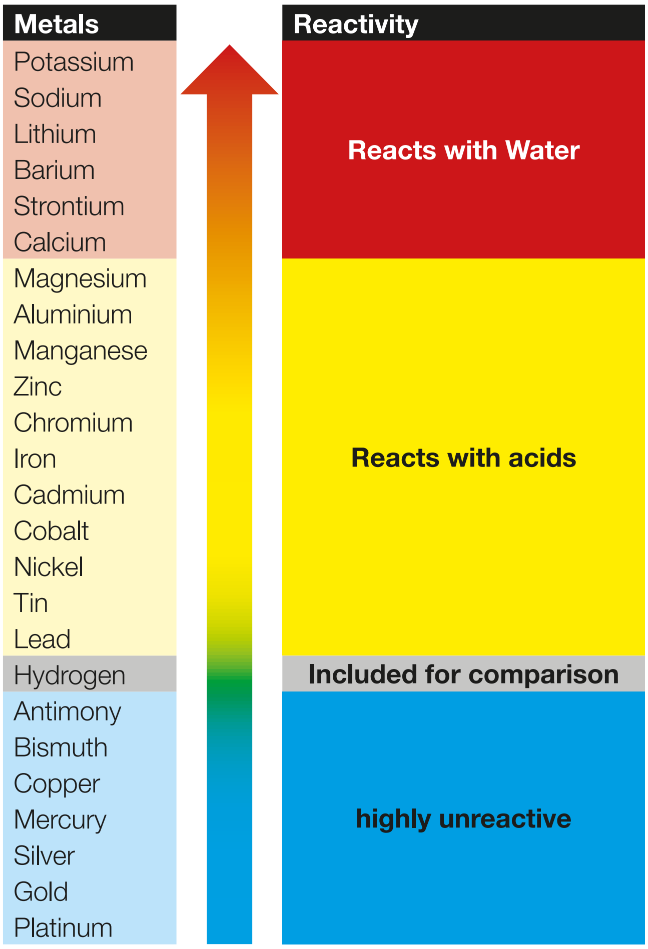

Figure 6.16: Metal reactivity determines which substances a metal will react with

Metals with the lowest reactivity (e.g., copper, mercury, gold, silver, platinum) require extreme conditions to undergo any reaction. For example, gold will not react with even the strongest and most concentrated mineral acids.

$$\mathrm{Au} + \mathrm{HNO_3} \rightarrow \text{no reaction}$$

Metals with medium reactivity (iron, tin, lead, etc.) will react with acids to release hydrogen but will not react with water.

$$\mathrm{Sn} + \mathrm{H_2O} \rightarrow \text{no reaction}$$

$$\mathrm{Sn} + 2\mathrm{HCl} \rightarrow \mathrm{SnCl_2} + \mathrm{H_2}$$

Reactions between metals and acids are analogous to metal displacement reactions, with the hydrogen ion taking the place of the metal cation as the electron acceptor. Remember that strong acids completely dissociate in solution: there are no 'HCl molecules', only separate H^+ ions and Cl^- ions. In the reaction of tin (Sn) and hydrochloric acid (HCl) shown above, the elemental tin donates its electrons to the hydrogen ions. The chloride ions are spectator ions.

$$\mathrm{Sn} + 2\mathrm{H^+} + 2\mathrm{Cl^-} \rightarrow \mathrm{Sn^{2+}} + \mathrm{H_2} + 2\mathrm{Cl^-}$$

Or, with the spectator chloride ions not shown:

$$\mathrm{Sn} + 2\mathrm{H^+} \rightarrow \mathrm{Sn^{2+}} + \mathrm{H_2}$$

Magnesium and aluminium are so reactive that they will react with steam or hot water to produce hydroxides and release hydrogen.

$$\mathrm{Mg} + 2\mathrm{H_2O} \rightarrow \mathrm{Mg(OH)_2} + \mathrm{H_2}$$

The most reactive metals (potassium, sodium, lithium, calcium) will react violently with cold water.

$$2\mathrm{Na} + 2\mathrm{H_2O} \rightarrow 2\mathrm{NaOH} + \mathrm{H_2}$$

Figure 6.17: Sodium metal reacting with water

In these reactions, the metal is so reactive that it is able to displace H^+ from the water molecule.

Metals with similar reactivities are ordered by comparing the *rates* of their reactions; more reactive metals will react more quickly. For example, both aluminium and zinc will react with steam to form hydrogen.

$$2Al + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2$$

$$Zn + 2H_2O \rightarrow Zn(OH)_2 + H_2$$

When the amount of hydrogen produced per second is compared for equivalent amounts of each metal, it is found that the aluminium reacts more quickly than the zinc. Hence, aluminium is more reactive than zinc.

Although it provides a very useful guide, there is no simple way to use a metal activity series to determine whether a particular metal-acid reaction will proceed. In fact, many reactions will occur under certain conditions (concentrated reactants, high temperature) but not under less favourable conditions. It is sufficient to be able to write a balanced equation for the reaction that might occur between any specific metal and acid (or water).

Science as a human endeavour:

Chemistry vs. Nazis

Despite being highly unreactive, gold **can** be dissolved in *aqua regia* ("king's water"), a 3-to-1 mixture of hydrochloric acid (HCl) and nitric acid (HNO_3). Although neither acid can dissolve gold alone, the mixture is effective in causing dissolution. Nitric acid, a powerful oxidiser, strips electrons from an extremely small number of gold atoms to produce gold cations (Au^{3+}). In pure nitric acid this process stops almost instantly, but the hydrochloric acid present in *aqua regia* reacts with the ions to produce tetrachloroaurate ions ($AuCl_4^-$). This removes the gold ions from solution and allows the oxidation of gold atoms by nitric acid to continue. In this way, gold is slowly dissolved.

Figure 6.18: Aqua regia containing partially dissolved gold

The dissolving power of *aqua regia* was known by George de Hevesy, a colleague of Niels Bohr, when the Nazis invaded Denmark in April 1940. Two German Nobel Prize winners—James Franck, a Jew, and Max von Laue, a Jewish sympathiser—had illegally sent their medallions to Bohr for safekeeping earlier in Hitler's reign, despite the Nazi ban on exporting gold from Germany. As Nazi troops entered Copenhagen, de Hevesy realised that the medallions were engraved with the names of the prize winners and that his team would be executed if they were discovered. He dissolved the medallions in aqua regia and placed the beaker containing the orange-coloured solution on a high shelf.

As a Jew himself, de Hevesy was forced to flee Denmark in 1943. When he returned after the Nazi's defeat, he found the beaker untouched, exactly where he had left it. He recovered the gold from solution as a precipitate and sent it to the Swedish Academy, which had originally awarded the Nobel Prizes. Using the very same gold from the original medallions, new medallions were cast and re-presented to Franck and von Laue in a ceremony in 1952.

Science understanding

Write equations and half-equations for reactions between a given acid and a nominated active metal.

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Reactions between metals and acids, or metals and water, are a type of redox reaction because they involve the transfer of electrons. Half-equations can be written by including only the species that is being oxidised or reduced, and then balancing the charge by adding electrons to one side of the equation.

Example

Write the oxidation and reduction half-equations for the reaction of magnesium and hot water to produce magnesium hydroxide and hydrogen.

First, identify the oxidation numbers of each atom.

$$\overset{0}{Mg} + 2\overset{+1}{H_2}\overset{-2}{O} \rightarrow \overset{+2}{Mg}(\overset{-2}{O}\overset{+1}{H})_2 + \overset{0}{H_2}$$

The oxidation number of magnesium has increased from 0 to +2. It has been oxidised.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

The oxidation number of one of the hydrogens in water has decreased from +1 to 0. It has been reduced. (The oxidation numbers of the oxygen and hydrogen atoms in the hydroxide ion are unchanged).

$$2H_2O + 2e^- \rightarrow 2OH^- + H_2$$

Example

Write the oxidation and reduction half-equations for the reaction of potassium and sulfuric acid.

First, identify the oxidation numbers of each atom.

$$2\overset{0}{K} + \overset{+1}{H_2}\overset{+6}{S}\overset{-2}{O_4} \rightarrow \overset{+1}{K_2}\overset{+6}{S}\overset{-2}{O_4} + \overset{0}{H_2}$$

The oxidation number of potassium has increased from 0 to +1. It has been oxidised.

$$K \rightarrow K^+ + e^-$$

The oxidation number of the hydrogens has decreased from +1 to 0. They have been reduced.

$$2H^+ + 2e^- \rightarrow H_2$$

The oxidation numbers of the atoms in the sulfate ions did not change because the sulfate ion did not participate in the reaction. It is a spectator ion, and it should not appear in either half-equation.

Questions

17. Write net ionic equations and fully balanced chemical equations for the metal displacement reactions below.

    (a) Zinc with tin(II) chloride ($SnCl_2$) solution. Zinc ions usually take a +2 charge.

    Net ionic equation:

    (1 mark) **KA1**

    Fully balanced chemical equation:

    (2 marks) **KA1**

(b) Strontium and copper(II) sulfate ($CuSO_4$) solution.

Net ionic equation:

(1 mark) **KA1**

Fully balanced chemical equation:

(2 marks) **KA1**

(c) Aluminium and nickel(II) nitrate ($Ni(NO_3)_2$) solution.

Net ionic equation:

(1 mark) **KA1**

Fully balanced chemical equation:

(2 marks) **KA1**

18. Manganese (Mn) dissolves in a solution of hydrochloric acid.

$$Mn_{(s)} + 2\ HCl_{(aq)} \rightarrow MnCl_{2(aq)} + H_{2(g)}$$

(a) Write a half-equation to show the oxidation that occurs in this reaction.

(2 marks) **KA1**

(b) Write a half-equation to show the reduction that occurs in this reaction.

(2 marks) **KA1**

19. Cobalt (Co) dissolves in a solution of phosphoric acid (H_3PO_4) to produce cobalt(II) phosphate ($Co_3(PO_4)_2$) and hydrogen (H_2).

(a) Write a half-equation to show the oxidation that occurs in this reaction.

(2 marks) **KA1**

(b) Write a half-equation to show the reduction that occurs in this reaction.

(2 marks) **KA1**

(c) Write a net ionic equation for the reaction.

(2 marks) **KA1**

(d) Write a fully balanced chemical equation for the reaction.

(2 marks) **KA1**

6.3: Electrochemistry

Science understanding

Electrochemical reactions involve a flow of electrons during a chemical reaction.

Galvanic cells produce electrical energy from spontaneous redox reactions.

Identify the anode and cathode and their charges, and the direction of ion and electron flow, in a galvanic cell, given sufficient information.

Write electrode half-equations for a galvanic cell, given sufficient information.

Metal displacement reactions occur spontaneously because they release energy. When these reactions occur directly between two substances, the energy that is released is converted mostly into unusable heat. However, by separating the two metals, the electrons can be forced to flow around a circuit. The flow of charge creates an electric field that can extract usable energy from the reaction. This concept is the basis of all electrical cells and batteries which turn chemical energy into electrical energy. Such cells are called **galvanic cells** after Luigi Galvani (or sometimes voltaic cells, after Alessandro Volta).

Consider the reaction of zinc metal and copper(II) sulfate. Zinc is more reactive than copper, so zinc will reduce the copper cations to produce copper metal.

$$Zn + CuSO_4 \rightarrow Cu + ZnSO_4$$

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

In a standard metal displacement reaction, the zinc is in direct contact with the Cu^{2+} solution. If they are instead connected by a conductive wire, electrons are forced to flow through the wire.

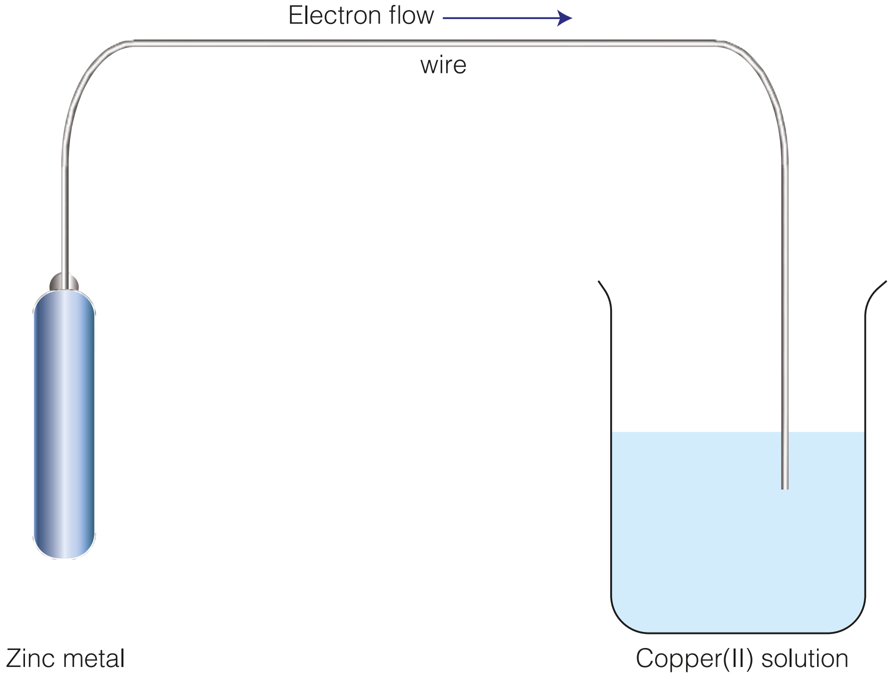


Figure 6.19: Separating the oxidising agent and the reducing agent forces the electrons to travel through a wire

However, in the situation shown above, the flow of electrons to the right would very quickly result in the build-up of negative charge in the copper solution. Negative charges repel one another, so this would prevent further electrons from flowing into the solution. Additionally, the Zn^{2+} cations have nowhere to go, causing a build-up of positive charge on the left. For both of these reasons, the reaction does not proceed.

To obtain a working galvanic cell, there must be a way for the charges to equilibrate. The remedy is to immerse the zinc metal in an electrolyte solution and include a **salt bridge**, a material soaked in saturated solution of an ionic compound, between the two electrolyte solutions.

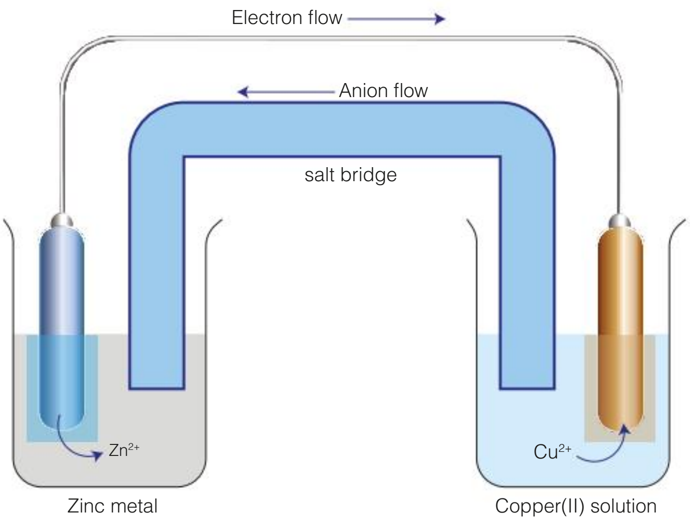


Figure 6.20: Adding a salt bridge allows the charges to equilibrate and completes the circuit

The bridge allows **anions** (negative ions) to flow to the electrolyte on the left, which contains newly produced Zn^{2+}, balancing the positive charge. It also allows **cations** (positive ions) to flow to the electrolyte on the right, replacing the Cu^{2+} ions which have been consumed.

An example of a functional galvanic cell is shown in Figure 6.21. The sample of zinc metal on the left is immersed in a solution of zinc(II) sulfate; this is the Zn/Zn^{2+} **half-cell**. It is connected with a wire to a sample of copper metal in a solution of copper(II) sulfate; this is the Cu/Cu^{2+} half-cell. The two solutions are also connected by the salt bridge, a piece of paper or fabric that has been soaked in sodium nitrate, $NaNO_3$.

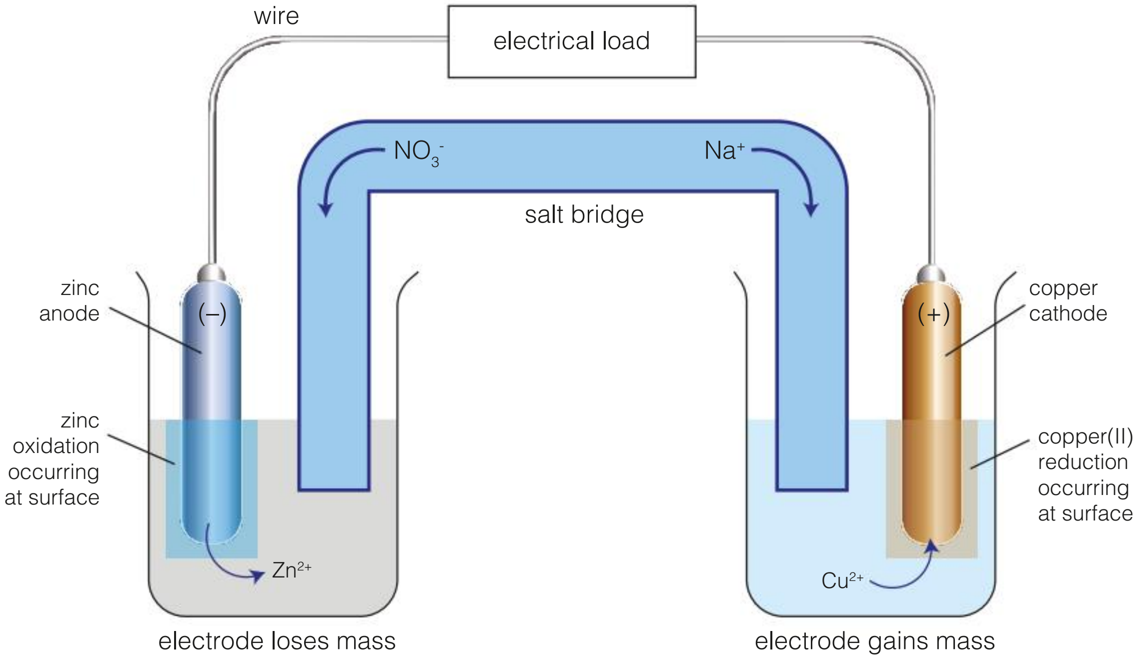


Figure 6.21: Functional galvanic cell

Zinc is more reactive than copper, so electrons flow through the wire from zinc to copper (left to right). An **electrical load**, for example a light bulb, can be included here to extract the energy that is released by the reaction. The electrode where oxidation occurs is called the **anode**. In this example, the anode is made of zinc. The anode is the source of the negatively charged electrons, so the anode is assigned a negative label.

$$Zn \rightarrow Zn^{2+} + 2e^-$$

The electrode where reduction occurs is called the **cathode**. In this example the cathode is made of copper (although it could be made from any conductive material, because the source of Cu^{2+} ions is the solution, not the electrode). Electrons flow toward the cathode, so the cathode is assigned a positive label.

$$Cu^{2+} + 2e^- \rightarrow Cu$$

A common way of remembering which process occurs at which electrode is to use the mnemonic RED CAT, which stands for '**RED**uction occurs at the **CAT**hode'.

Galvanic Cell	Anode	Cathode
Process	oxidation	reduction
Charge	negative	positive
Reactivity	more reactive metal	less reactive metal
Energy Conversion	chemical to electrical	

Science as a human endeavour:

Sacrificial Anodes

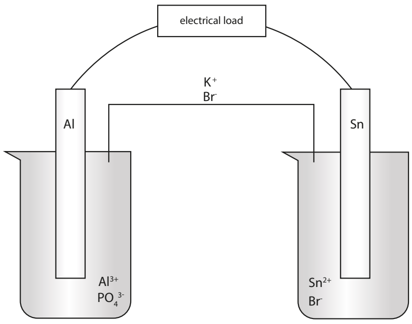
When metal surfaces are submerged in electrolytes, they can undergo a process called **corrosion**. The metal atoms donate electrons to more electronegative substances in the electrolyte and return to anionic form. The most well-known example of corrosion is **rusting**, which involves atoms of Fe donating electrons to water molecules and forming iron(II) oxide, Fe_2O_3, which gives rust its characteristic reddish-brown colour.

Many important structures that are made from iron must come into contact with electrolytes, including pipelines, tanks, and the hulls of ships. These structures are often protected with the use of a **sacrificial anode**, a sample of a more reactive metal that preferentially reacts with the ions in electrolytes. For example, small plates of zinc attached to the hulls of ships corrode instead of the iron hull, maintaining the integrity of the hull.

Figure 6.3.4: Partially corroded sacrificial zinc anode

Questions

20. A galvanic cell is constructed using an aluminium electrode in a solution of aluminium phosphate (the Al/Al^{3+} half-cell) and a tin electrode in tin(II) bromide (the Sn/Sn^{2+} half-cell). Potassium bromide (KBr) is used as the electrolyte in the salt bridge.



(a) State whether aluminium or tin is more reactive.

(1 marks) **KA1**

(b) On the diagram, label the direction of electron flow on the galvanic cell.

(1 mark) **KA1**

(c) State whether the aluminium is oxidised or reduced.

(1 mark) **KA1**

(d) On the diagram, label the anode.

(1 mark) **KA4**

(e) On the diagram, label each electrode as either positive (+) or negative (–).

(1 mark) **KA4**

(f) Write the half-equation for the oxidation of atoms at the anode.

(2 marks) **KA1**

(g) Write the half-equation for the reduction of ions at the cathode.

(2 marks) **KA1**

(h) Write the overall cell reaction by combining the half-equations. Show your working. You do not need to show spectator ions.

(3 marks) **KA1**

(i) The salt bridge in this cell is a strip of paper soaked in potassium bromide solution. State one function of the salt bridge in the cell.

(1 mark) **KA1**

(j) On the diagram, label the direction of the flow of potassium ions in the salt bridge.

(1 mark) **KA1**

Science understanding

Draw a diagram of a galvanic cell, given sufficient information.

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A diagram of a galvanic cell should include two electrodes each immersed in an electrolyte solution. The electrodes should be connected with a conducting wire across an electrical load, and the electrolytes should be connected with a salt bridge.

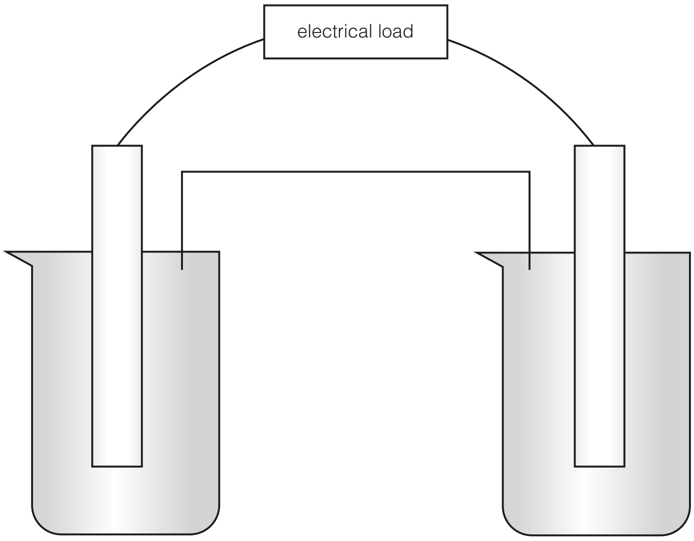

Figure 6.23: Unlabelled diagram of a galvanic cell

As a minimum, the diagram should include labels showing:

- which electrode is the anode and which is the cathode
- the substance the anode and cathode are made from
- the anode is assigned as 'negative'
- the cathode is assigned as 'positive'
- the direction of electron flow
- the direction of anion and cation flow in the salt bridge

If the electrolytes used in each half-cell and in the salt bridge are known, these can also be included in the diagram.

The diagram below shows a galvanic cell made from iron and lead. Iron nitrate is used as the **anodic** electrolyte, lead nitrate is used as the **cathodic** electrolyte, and potassium nitrate is used in the salt bridge.

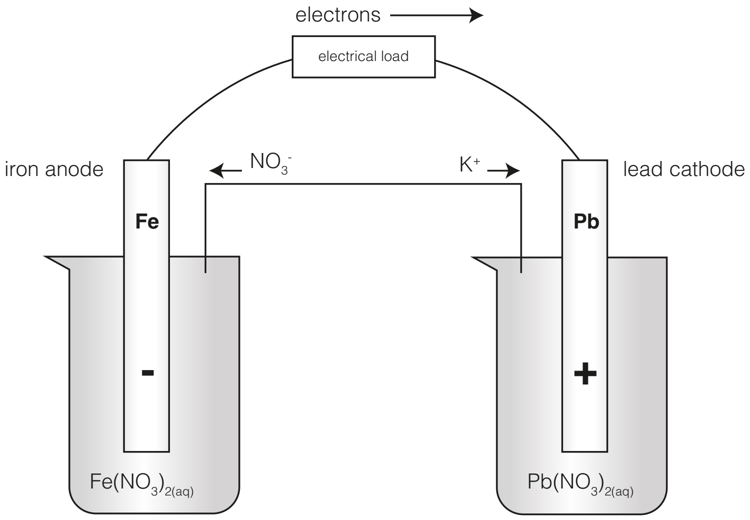

Figure 6.24: Labelled diagram of a galvanic cell

If the more reactive metal is placed on the left, then this will always be the anode and electrons will always flow from left to right.

Questions

21. A galvanic cell was constructed using a nickel electrode in nickel(II) sulfate solution ($NiSO_4$) and a copper electrode in copper nitrate solution ($Cu(NO_3)_2$). A strip of paper soaked in copper(II) sulfate ($CuSO_4$) was used as the salt bridge.

(a) Draw and label a diagram of the cell.

(6 marks) **KA4**

(b) Name the metal that is oxidised in the cell.

(1 mark) **KA1**

(c) Write a half-equation to show the reaction at the anode in the cell.

(2 marks) **KA1**

(d) Write a half-equation to show the reaction at the cathode in the cell.

(2 marks) **KA1**

(e) Combine the half-equations to write the net ionic reaction.

(2 marks) **KA1**

Science understanding

Galvanic cells are commonly used as portable sources of electric current.

Compare the operation of different types of batteries.

The voltage produced by a galvanic cell depends on the potential energy of the electrons in the two metals used, which can be inferred from the metal reactivity series. For example, a galvanic cell made from zinc and copper is observed to produce 1.10 V because zinc is significantly more reactive than copper. However, a galvanic cell made from zinc and cobalt produces only 0.59 V because, while zinc is more reactive than cobalt, the difference in reactivity is smaller. To produce higher voltages, galvanic cells can be connected in series to form **batteries**.

One of the most common sources of portable electrical energy is alkaline "batteries". Despite the name, these are single electrochemical cells which use zinc as the anode and manganese dioxide (MnO_2) as the cathode. An alkaline solution, generally aqueous potassium hydroxide (KOH), permeates the inside of the cell and is used as the electrolyte. A polymer that conducts ions and is stable in very alkaline conditions is used as the salt bridge.

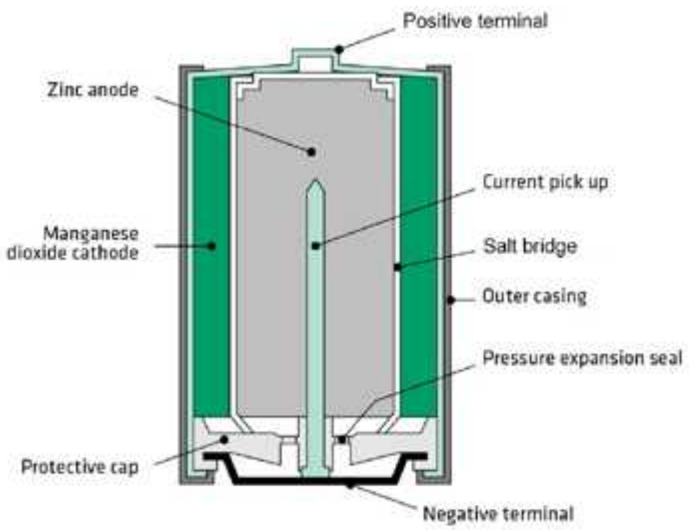

Figure 6.25: Schematic diagram of Zn/MnO_2 alkaline "battery"

When the cell is in operation, zinc is oxidised by the hydroxide ions at the anode to produce water and free electrons.

$$\overset{0}{Zn}_{(s)} + 2\overset{-2}{O}\overset{+1}{H}^-_{(aq)} \rightarrow \overset{+2}{Zn}\overset{-2}{O}_{(s)} + \overset{+1}{H_2}\overset{-2}{O}_{(l)} + 2e^-$$

The electrons flow through the negative terminal and into the device that is being powered by the cell. As they move, they generate an electric field which converts the chemical potential energy into whatever useful form of energy is desired (e.g., electromagnetic energy from a light bulb, kinetic energy from a motor). The cell is being **discharged**.

Simultaneously, electrons flow from the external device into the positive terminal and to the cathode, where they combine with water and manganese dioxide to produce manganese trioxide (Mn_2O_3) and hydroxide ions.

$$2\overset{+4}{Mn}\overset{-2}{O}_{2(s)} + \overset{+1}{H_2}\overset{-2}{O}_{(l)} + 2e^- \rightarrow \overset{+3}{Mn_2}\overset{-2}{O}_{3(s)} + 2\overset{-2}{O}\overset{+1}{H}^-_{(aq)}$$

Equal amounts of hydroxide are consumed and produced in the two half-equations, and potassium is a spectator ion, so the amount of electrolyte remains constant. The overall reaction that occurs is:

$$\overset{0}{Zn}_{(s)} + 2\overset{+4}{Mn}\overset{-2}{O}_{2(s)} \rightarrow \overset{+2}{Zn}\overset{-2}{O}_{(s)} + \overset{+3}{Mn_2}\overset{-2}{O}_{3(s)}$$

Eventually, the zinc metal is consumed, and the reaction stops. When this happens, the battery has gone flat. Galvanic cells which utilise non-reversible redox reactions are used once and then discarded. These are called **primary cells**, and an alkaline "battery" is one example.

Secondary cells function similarly but, because they use reversible redox reactions, they can be **recharged**. During recharge, electrical energy from some other source is applied to the system and forces the flow of electrons to reverse. This causes the reverse redox reaction to occur: the less reactive metal can donate its electrons to the more reactive metal using the external energy that is supplied. In this way, electrical energy is converted back into chemical potential energy.

During recharge of a galvanic cell, the identities of the electrodes are flipped because the flow of electrons has been reversed. The more reactive metal becomes the positive cathode, and the less reactive metal becomes the negative anode. Whether the cell is discharging or recharging, oxidation always occurs at the anode, and reduction always occurs at the cathode.

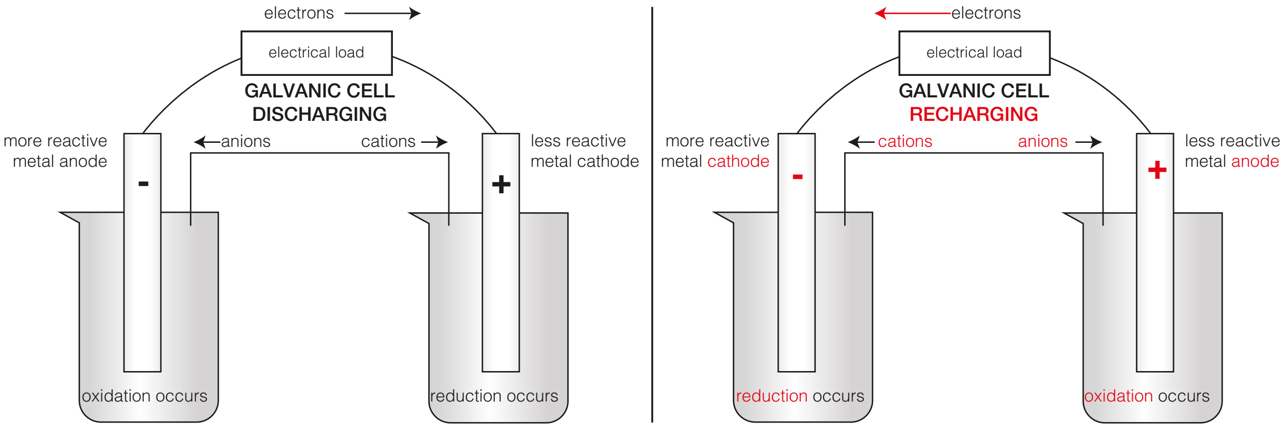

Figure 6.26: Galvanic cell during discharge and recharge

Galvanic Cell Discharge	Anode	Cathode
Process	oxidation	reduction
Charge	negative	positive
Reactivity	more reactive metal	less reactive metal
Energy Conversion	chemical to electrical	

Galvanic Cell Recharge	Anode	Cathode
Process	oxidation	reduction
Charge	**positive**	**negative**
Reactivity	**less reactive metal**	**more reactive metal**
Energy Conversion	**electrical to chemical**	

Science as a human endeavour:

South Australia's Big Battery

The increasing amount of renewable energy generation worldwide has created demand for high-capacity energy storage on electricity grids. Lithium-ion batteries have been used in small devices since the 1990s but, since 2017, large-scale lithium-ion batteries have begun to compete with other energy-storage technologies in real-world applications.

In response to the widespread power outage in South Australia in September 2016, which affected nearly the entire state for more than three hours, the SA Government contracted with Tesla, Inc. to build the Hornsdale Power Reserve (HPR). At the time of installation, it was the largest and most powerful electrochemical battery system ever built, capable of storing 129 MWh of energy and delivering 100 MW of power.

The HPR stabilises the grid by responding to fluctuations in power generation in other parts of the network much faster than alternative technologies (e.g., gas-fired power). It is estimated to have saved grid operators more than $100 million in 2019, some portion of which will have been passed onto consumers as reduced electricity bills.

Questions

22. Lead-acid batteries are used to provide the spark that ignites combustion in motor vehicles. Lead metal and lead compounds are used as the electrodes and are immersed in sulfuric acid as the electrolyte. Most car batteries use six 2V cells connected in series to produce a total of 12 V.

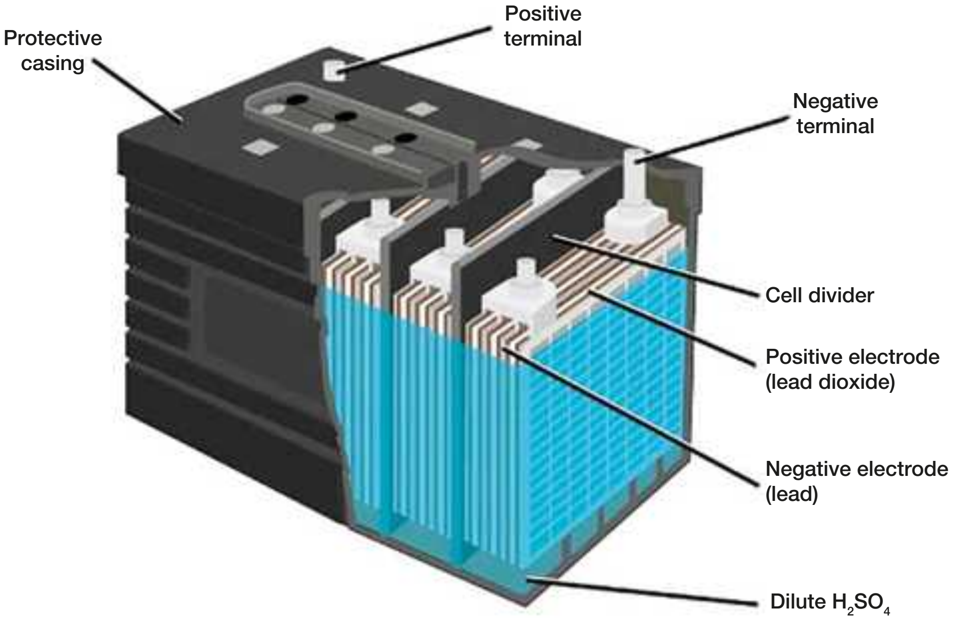

Figure 6.27: Lead-acid battery schematic

In one half-cell, the following reaction occurs during discharge:

$$Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$$

(a) Identify the change in oxidation number of the lead.

.... *(1 mark)* **KA1**

(b) Hence, state whether lead has been oxidised or reduced.

.... *(1 mark)* **KA1**

(c) Hence, state whether lead metal serves as the anode or the cathode during discharge.

.... *(1 mark)* **KA4**

At the other electrode, lead(IV) oxide reacts with sulfuric acid to produce lead(II) sulfate and water.

(d) Balance the following half-equation showing the reduction of lead(IV) oxide to form lead(II) sulfate in acidic conditions.

$$PbO_2 + SO_4^{2-} \rightarrow PbSO_4$$

(3 marks) **KA1**

(e) Lead(II) sulfate is not soluble in aqueous conditions. State whether the sulfate ions are spectator ions.

.... *(1 mark)* **KA2**

(f) Explain your answer to Question 22 (e).

.... *(1 mark)* **KA2**

Summary Test 6: Redox Reactions

1. When powdered potassium carbonate is placed in nitric acid, effervescence is observed. The reaction produces potassium nitrate, water, and carbon dioxide.

$$K_2CO_3 + 2HNO_3 \rightarrow 2KNO_3 + H_2O + CO_2$$

(a) Determine the oxidation number of each atom in the reaction and write these on the equation above.

(5 marks) **KA1**

(b) Define oxidation in terms of a change in oxidation number.

(1 mark) **KA4**

(c) Hence, state and explain whether the reaction is a redox reaction.

(2 marks) **KA1**

2. Calcium metal reacts with gaseous fluorine to produce calcium fluoride.

$$Ca_{(s)} + F_{2(g)} \rightarrow CaF_{2(s)}$$

(a) State the oxidation number of fluorine before the reaction occurs.

(1 mark) **KA1**

(b) State the oxidation number of fluorine after the reaction occurs.

(1 mark) **KA1**

(c) Write a half-equation to show the reduction of fluorine in the reaction.

(2 marks) **KA1**

(d) State the number of electrons gained by each fluorine *atom* in the reaction.

(1 mark) **KA1**

(e) Hence, state the number of electrons accepted by each fluorine *molecule* in the reaction.

(1 mark) **KA2**

(f) Write a half-equation to show the oxidation of calcium in the reaction.

(2 marks) **KA1**

3. When zinc metal is mixed with solid iodine in water, the iodine partially dissolves and a reaction occurs producing zinc iodide.

$$Zn_{(s)} + I_{2(aq)} \rightarrow ZnI_{2(aq)}$$

(a) Write a half-equation to show the oxidation of zinc.

(2 marks) **KA1**

(b) Write a half-equation to show the reduction of iodine.

(2 marks) **KA1**

(c) State whether zinc or iodine is the oxidising agent.

(1 mark) **KA1**

4. Nitric acid is extremely corrosive. As copper dissolves in concentrated nitric acid, it produces nitrogen monoxide (common name: nitric oxide).

Balance the following chemical equation by first writing half-equations.

$$Cu + HNO_3 \rightarrow Cu^{2+} + NO$$

Balanced chemical equation:

(6 marks) **KA1**

5. A strip of magnesium metal (Mg) is placed in a solution of copper(II) sulfate ($CuSO_4$). A coating of copper metal forms on the magnesium strip.

(a) Write a half-equation to show the oxidation that is occurring in this reaction.

(2 marks) **KA1**

(b) Write a half-equation to show the reduction that is occurring in this reaction.

(2 marks) **KA1**

(c) Write a balanced chemical equation for the overall reaction.

(2 marks) **KA1**

(d) Identify the reducing agent in the reaction.

(1 mark) **KA1**

(e) Explain why reactions such as these are called ‘displacement’ reactions.

(2 marks) **KA4**

6. With reference to the metal activity series shown, explain why nuggets of pure gold can be found in nature, while sodium can only be stored safely under an inert substance such as mineral oil.

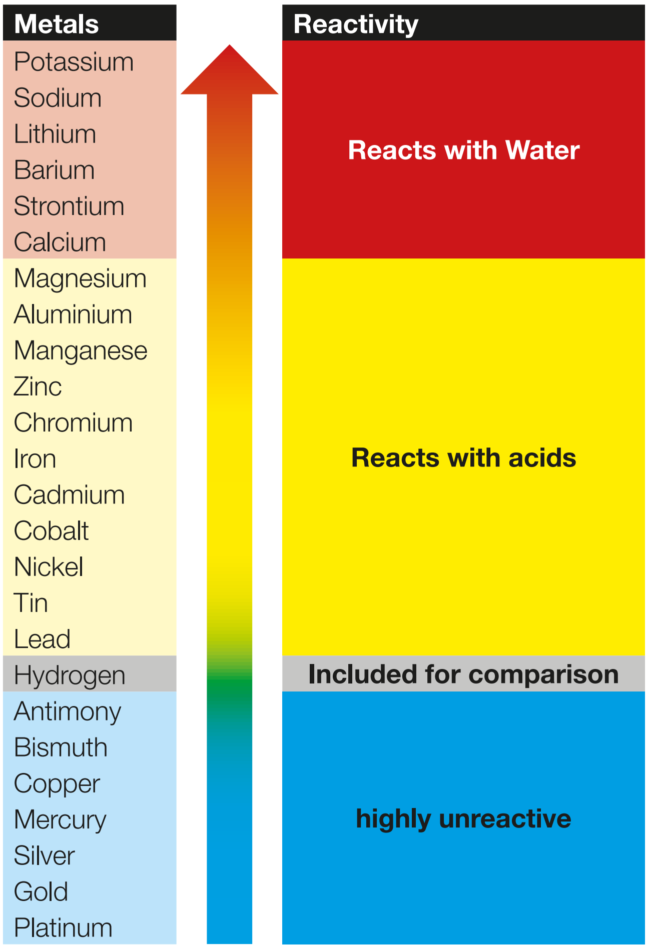

Figure 6.28: Metal activity series

(2 marks) **KA2**

7. A galvanic cell was constructed using a manganese electrode in manganese(II) chloride solution ($MnCl_2$) and a lead electrode in lead(II) nitrate solution $Pb(NO_3)_2$. A piece of fabric soaked in sodium chloride (NaCl) was used as the salt bridge.

Draw and label a diagram of the cell.

(6 marks) **KA4**

8. Lithium-ion batteries typically use graphite and cobalt(IV) oxide as electrodes. During discharge, the following reaction occurs at the cobalt(IV) oxide electrode.

$$CoO_2 + Li^+ + e^- \rightarrow LiCoO_2$$

(a) Determine the oxidation number of each atom in the reaction.

.... *(3 marks)* **KA1**

(b) Hence, state which atom has gained an electron.

.... *(1 mark)* **KA1**

(c) State whether this reaction occurs at the anode or cathode.

.... *(1 mark)* **KA1**

Topic 7: Science inquiry skills

All disciplines of science involve the process of scientific inquiry. Inquiry is a process of extracting information from a source by asking questions. Scientific inquiry over the course of history has allowed for the development of our society.

Science as a human endeavour

History of scientific inquiry

The history of scientific inquiry is not a linear one and involves contributions from a number of significant time periods, and civilisations. The Greeks were the first in western civilisation to include observation and measurement in their learning about the world. However, there was not enough structure to define this as the scientific process. Aristotle was the founder of empirical science (science which is observable and based on evidence); however, many historians regard science as starting during the golden age of Islam between the 10th and 14th century. During this period, Islamic scholars developed scientific processes which began to resemble our modern process. The scientific processes created during this time became the basis for western science during the Renaissance. Scholars during this time insisted on repeatability, and introduced ideas such as peer review. Empirical science was refined during the Renaissance, and impressive thinkers of this time such as Rene Descartes proposed that deduction was imperative to learning and forming an understanding of the world.

Over the following centuries, the world was exposed to the innovation and genius of Isaac Newton, Albert Einstein, Bertrand Russell, Karl Popper and many others who further refined the scientific process.

Scientific method

In chemistry, investigation is an integral part of the learning and understanding of concepts. This uses the scientific method to test ideas and develop new knowledge.

Scientific method is an organised and systematic approach to investigating a variable using laboratory techniques. This allows data collection, the results of which can be then analysed to answer a particular question. In other words, it is a process involving a series of steps that allow chemists to qualitatively and/or quantitatively investigate facets of the discipline.

An outline of each step is presented below for one scientific method.

Question or Aim

For any investigation to be conducted, first a chemist must present a problem or concept that they want to explore. This must be transformed into an experimental question or aim which effectively provides the purpose of the experiment. If presented as an aim, it commonly starts with “to investigate”.

Background

With the purpose of the investigation clearly defined, a chemist must conduct background research. This may include reading scientific reports from similar experiments, preliminary reading regarding the techniques required, or building a better understanding of the chemistry for the investigation being undertaken.

Hypothesis

Once an experiment has a distinct and defined purpose that has been researched, the chemist sometimes state the hypothesis. A hypothesis is an idea that could explain an observation. It can be tested by performing an investigation. It must be testable within the laboratory it is to be conducted in to be considered a valid and sound hypothesis. A well-written hypothesis should refer to both the independent and dependent variable.

The hypothesis often predicts the suggested trend of the data that will be obtained. A single experiment provides evidence to support or not support the stated hypothesis.

Variables

From the development of hypothesis, the variables can be defined. Variables are factors which can be changed, measured or controlled during an experiment. The types of variables will be explored in Topic 7.2.

Experiment/Investigation

Following the development of the preliminary facets of the investigation, the proposed experimental procedure is conducted in the laboratory or in the field. In some instances, the experimental method may not be productive in producing observable or expected data. In these instances, refinement of the procedure may be required and a second experiment conducted.

Observations and/or measurements taken during the experiment will have been recorded using an appropriate format, usually a table.

Data analysis

Data obtained from the experiment must then be analysed. In chemistry, this may involve using the collected data to construct a graph, undertaking calculations, or referring to previously known data to allow conclusions to be formed from the chemist's data. In essence, this should define the relationship between the independent and dependent variables, and lead to a response to the investigation question. However, if the experimental data does not lead to this, refinement of the hypothesis may once again be required and an additional experimental procedure carried out.

The experimental data provides evidence to support or not support the stated hypothesis. Alternatively, the evidence may provide inconclusive evidence to support or reject the stated hypothesis. There is never "proof" from conducting an experiment once.

Conclusions

With the data analysed, a clear conclusion should be able to be determined. This should refer directly to the aim and hypothesis, refer to the variables, and communicate the relationship between the independent and dependent variables, as demonstrated by the data collected.

Such conclusions may direct the aim of future investigations to explore concepts further.

Communication of results

All components of scientific method must be adequately communicated in the appropriate form of a scientific report.

Details of this are outlined throughout this chapter.

The steps of the scientific method are summarised in Figure 7.01.

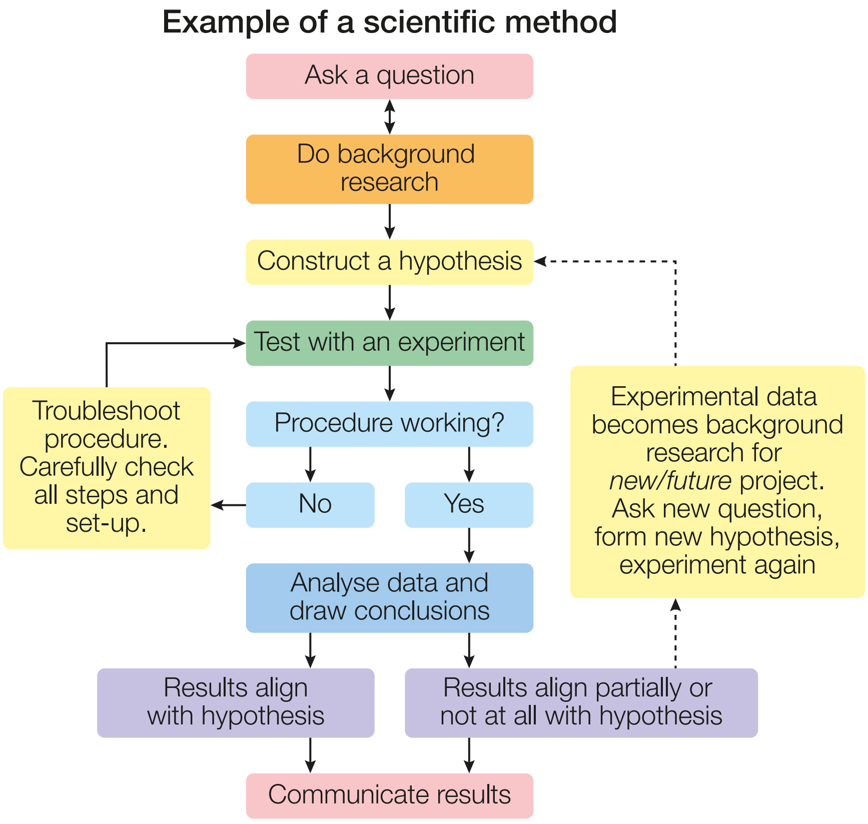

Figure 7.01: The main steps of the scientific method.

Questions

1. Shane noticed that his favourite pants were not as clean as they used to be. His friend Brad suggested he should try Cleaning Prime detergent, a new laundry product found at the supermarket.

   Shane made sure to wash one pair of pants in tap water, another in his original detergent and a final pair in the new Cleaning Prime detergent. After repeating this three times for each detergent, the pants washed in Cleaning Prime detergent did not appear to be any cleaner than the pants washed in tap water or his original detergent.

   (a) State the problem that initiated Shane's investigation.

   (1 mark) **KA1**

   (b) Use Shane's problem above to describe how he implemented the scientific method.

   (3 marks) **KA2**

   (c) From the information provided what is the likely conclusion that would be drawn from Shane's investigation.

   (1 mark) **KA2**

   (d) Propose if a sound hypothesis can be written for Shane's problem. Explain why or why not.

   (2 marks) **KA2**

2. Identify the following hypotheses that are testable in a laboratory :

Hypothesis	Testable (Y/N)
Gold has a more attractive lustre than nickel	
The heat of combustion of an alcohol increases as the length of the carbon chain of the alcohol increases	
The rate of reaction between a metal and an acid increases as the temperature of the acid increases	
Nitrogen fertilisers are best for plants	

(4 marks) **KA4**

Investigation procedures (design and implementation)

Scientific understanding

Scientific methods enable the systematic investigation to obtain measurable evidence.

Obtaining meaningful data depends on conducting investigations using appropriate procedures and safe, ethical working practices.

Application of the scientific method allows the implementation of chemical investigations in a laboratory.

An important part of the scientific method is designing well-rounded investigations which will produce data that will allow the formation of a valid conclusion, and the subsequent communication of these findings in a scientific report. Further aspects of the below components will be explored throughout this chapter.

A scientifically valid investigation should include:

Question or Aim

A statement which explains the purpose of the investigation.

Hypothesis

An educated inference that refers to the variables being investigated.

Identification of variables

As introduced earlier, variables are factors which can change in an experiment.

- Independent variable

  The independent variable is the factor that is changed in the experiment. For an experiment to be valid, only one independent variable must be changed at a time.

- Dependent variable

  The dependent variable is the factor of the experiment which changes in response to variations in the independent variable.

- Controlled variables

  Controlled variables are quantities which must be held constant to ensure that any change to the dependent variable is a result of changes to the independent variable. This is imperative to ensure that other factors are not affecting the outcome of the investigation. Scientists must define how and why these variables are controlled.

Example

Consider an investigation titled:

Investigating the rate of carbon dioxide production when marble chips are reacted with varying concentrations of hydrochloric acid.

The independent variable is the *concentration of hydrochloric acid*, as it is the variable which is altered through the course of the investigation.

The dependent variable is the *rate of carbon dioxide production*, as it is the variable which is measured during the investigation.

Controlled variables could include:

- volume of hydrochloric acid
- source of hydrochloric acid
- cannot control extraneous
- time allocated for each reaction trial.

These variables must be maintained consistently, as failure to do so will introduce variance in the results obtained, subsequently introducing sources of errors.

Materials and apparatus list

All investigations must include a detailed list of all materials (reagents) and apparatus (laboratory equipment) that are required for the reader to reproduce the experiment. When selecting equipment, you need to consider the resolution of the measuring instrument. Some variables need to be measured using a finer scale to make the results clear.

The materials and apparatus list should be inclusive of details such as:

- quantities required
- concentrations of solutions
- volume of measuring equipment or glassware used.

Procedure

Note: It is always important to remember that a procedure should be detailed and clear enough that a second party can replicate the same experiment from the described steps.

A procedure or method must be designed in a way which will allow the measurement of the dependent variable. Procedures are developed by:

- conducting research about the background chemistry and previous similar experiments
- being aware of the materials and apparatus available in the laboratory being used
- ensuring that written steps are numbered, are logically communicated and are detailed in their articulation.

Ethical and safety considerations

Through the development of the procedure and materials list, ethical and safety considerations should be identified.

- **Ethics:** Ethical standards in chemistry promote norms in the research being conducted. They prohibit the fabrication or falsification of data to uphold the reliability of data. This ensures that the experimenter is accountable for the investigation being undertaken.

  The concept of experimental (or research) bias is also important to consider in discussion of ethics. This occurs when the experimental results are influenced in order to portray a particular outcome. This can emerge from the presence of certain errors, or from incomplete consideration of the effect of all variables. Bias can also be introduced by the selection of particular samples, subjects or materials that are more likely to generate desired results.

  Ethics referring to maintenance of moral or social values are also valid to consider, to ensure that investigations are in compliance with the law and public health and safety.

- **Safety:** A completed risk assessment should accompany every chemical investigation. A risk assessment is the determination of the quantitative or qualitative risk related to a practical activity. It will outline any potential hazards of the chemical reactants, products, and equipment being used.

  A well-constructed risk assessment allows the experimenter to be educated about possible hazards before undertaking the investigation, and therefore ensures that correct safety precautions such as PPE (personal protective equipment) are applied.

Experimental results

Any measurement or observation of the dependent variable must be recorded throughout the course of an investigation using an appropriate format. This will be inclusive of raw and processed data.

Graphs produced from data are also included as part of the results of the investigation, as the constructed graphs will be used in the analysis to formulate trends and predictions about the data. Calculations may be required to determine derived data or to finalise required parameters of an investigation.

Analysis of data

A crucial part of any investigation is the analysis of the collected results. This should include statements regarding the trends of graphs, discussion of the calculated parameters, and observations of the experiment. Any discussion of results should be inclusive of actual data values and may also compare the collected data to known true or literature values of data.

Critical evaluation of procedures and outcomes

In addition to the analysis of results, the investigation design is reviewed. This should be inclusive of a discussion of random and systematic errors, precision and accuracy of data. Possible improvements to the experimental design which could be incorporated the next time the experiment is performed is also beneficial.

Conclusion

The conclusion of the investigation should reflect the stated aim of the experiment, communicating a response to the question and aim. In addition to this, it should state if the hypothesis of the experiment was supported or not supported by the data collected, and, briefly, connect this to the supporting chemistry of the investigation.

Individual and collaborative work

Experimenters are required to work both individually and collaboratively, depending on the type of research being conducted. When individual work is undertaken, experimenters must be able to demonstrate initiative, solve problems and make decisions independently. They must be able to manage their time, demonstrate an ability to complete all techniques required by the method, and communicate all findings effectively.

As modern research methods have become more specialised, and the complexity of investigations has amplified, collaboration among experimenters trained in the same or different fields have become increasingly prevalent. Collaborative work involving two or more experimenters will involve each bringing expertise to assess the investigation question or aim. They may work on separate parts of the investigation, including the procedure design, operation of equipment, observation of reactions or analysis of data. Each member of the team must demonstrate an ability to provide feedback to others, and execute self-reflection on their performance and contributions. They must share responsibility, communicate with one another effectively, demonstrating a good work ethic at all times.

The outcomes of an investigation (inclusive of all of the above criteria) are always shared, and thus their communication in a report is paramount, independent of whether the investigation was conducted individually or collaboratively. This will often (if not always) involve a peer review, where others from the same discipline of science evaluate the presented work.

Questions

3. Read the below passage about cow's milk.

Cow's milk is an excellent source of calcium for human nutrition. In general, the gross composition of cow's milk is 87.7% water, 4.9% lactose (carbohydrate), 3.4% fat, 3.3% protein, and 0.7% minerals. The pH of milk is 6.3 - 6.6, which allows the protein molecules to be dispersed evenly in the solution.

When milk is acidified (commonly done with vinegar), it separates into what is known as curds and whey. The solid curds are comprised of a type of protein called casein. When sodium bicarbonate is added to the solid curds, neutralisation occurs and a liquid is formed again. Liquid casein is a natural glue. There have been many uses for casein glue throughout history. Records show it was created and used by the ancient Egyptians. It was employed in the Middle Ages to bind together thin panels into thicker plates on which artists could paint.

The glue consists of molecules of the casein protein that are precipitated from the milk by adding the acid. It is the polymerisation of these protein molecules that forms a functional glue.

Interestingly, the fat present in milk can interfere with the polymer chains, lubricating them like oil does in a bicycle chain, preventing them from sticking together as effectively.

(a) From the above passage, identify a factor could be changed to form the basis of an experiment investigating how the most effective glue can be made from milk.

(1 mark) **KA1**

(b) State the name given to the variable changed during an investigation.

(1 mark) **KA1**

(c) Suggest a possible aim for this experiment.

(1 mark) **KA1**

(d) Propose a sound hypothesis for the investigation.

(2 marks) **KA1**

(e) Describe how the efficacy of the glue be tested in a laboratory setting.

(2 marks) **KA1**

(f) From the information given, suggest two variables that will need to be controlled over the course of the investigation.

(2 marks) **KA1**

(g) Suggest what background chemistry should be researched to allow a broad understanding of what is being investigated.

(2 marks) **KA1**

4. A student carried out the following experiment:

She added equal masses of glucose to three beakers containing equal volumes of hot water. These beakers were cooled to three different temperatures. The mass of sugar which crystallised in each beaker was then measured.

(a) State the independent variable of the experiment.

(1 mark) **KA1**

(b) State the dependent variable of the experiment.

(1 mark) **KA1**

(c) Write a implied hypothesis for this experiment.

(2 marks) **KA1**

5. A student has designed an experiment to determine the relationship between the mass of powdered zinc and the volume of hydrogen produced when it is added to a solution of dilute hydrochloric acid.

(a) Write a suitable hypothesis for the experiment.

(2 marks) **KA1**

(b) Identify the independent variable.

(1 mark) **KA1**

(c) Identify the dependent variable.

(1 mark) **KA1**

(d) Identify three factors that need to be controlled in this experiment.

(1 mark) **KA1**

(e) Describe one aspect of safety that would need to be considered in this practical. Include how to minimise any presented risk.

(2 marks) **KA2**

(f) The student used different lengths of zinc metal, and for each length carried out three trials. Explain why this was done.

....

....

.... *(2 marks)* **KA2**

(g) In the practical report, the student stated that by completing three trials of each length of zinc metal the accuracy of the result increased. Comment on the student's statement.

....

....

.... *(2 marks)* **KA2**

(h) If this practical was carried out in teams of two, describe what constructive and productive collaboration would look and sound like?

....

....

.... *(2 marks)* **KA2**

Presenting results in investigations

Scientific understanding

Results of investigations are presented in a well-organised way to allow them to be interpreted.

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In addition to the above components of a scientific investigation, an experimenter should also consider the way in which investigation results are presented and communicated. This includes the use of appropriate units, and the construction of effective tables and graphs.

The International System of Units (SI)

Much of the chemistry conducted in a laboratory will have quantitative aspects to it. Therefore, the correct application of units is imperative. This may include the measurement of a dependent variable, observation of the room temperature to ensure that it remains constant, or calculation of a solution's concentration after a titration. Such measurements need to be able to be recorded and consistently understood by others in the scientific community when they are communicated in a report.

SI is the metric system used for investigations and is the globally accepted scientific unit system. A standard for scientific measurements ensures that data is accurate, precise, reproducible and unchanging. Within the SI there are seven fundamental (base) units, from which all other units are derived. They are the:

- Metre (m), for length
- Kilogram (kg), for mass
- Second (s), for time
- Kelvin (K), for temperature
- Ampere (A), for electric current
- Candela (cd), for luminosity
- Mole (mol), for the amount of substance

Derived units are subsequently calculated from fundamental units and include units such as:

- Volume ($m \times m \times m = m^3$) or
- Celsius temperature °C = °K (0 °K = –273°C)

Note: All units use a prefix (related to the fact that the metric system was built on powers of ten). These prefixes are useful when converting from or to an SI unit.

Table 7.01 demonstrates common prefixes useful for unit conversion.

Symbol	Prefix	Multiplication factor
T	Tera	10^{12}
G	Giga	10^{9}
M	Mega	10^{6}
k	Kilo	10^{3}
C	Centi	10^{-2}
m	Milli	10^{-3}
µ	Micro	10^{-6}
n	Nano	10^{-9}
p	Pico	10^{-12}

Table 7.01: Unit prefixes

Science as a human endeavour

Metric system

The metric system was introduced in 1791 during the French Revolution and was adopted as the standard in 1875. In fact, the history of the 'metre' can be traced back to the 18th century. At that time, there were two theories regarding the manner in which a standard unit of length was defined. The 'pendulum' model posed that the metre was the length of a pendulum having a half-period of one second. The 'meridian' model defined the metre as one ten-millionth of the length of the Earth's meridian along a quadrant (one quarter of the Earth's circumference).

It was in 1791 that the French Academy of Sciences verified the meridian model as the accepted standard. This was chosen over the pendulum model, because the force of gravity varies slightly over the surface of the Earth and this could subsequently affect the period of a pendulum.

The introduction of the metric system was imperative, as up until this time there were around 250,000 different units of measure in France alone. All seven fundamental units of measure were accompanied by a physical prototype. For weight, this was, and is still known as Le Grand K or the International Prototype Kilogram. It is well protected in a vault near Paris, and is actually vacuum-sealed under three bell jars. Le Grand K was forged in 1897 and is a palm-sized metal cylinder composed from an alloy of platinum and iridium. It is commonly referred to as the perfect kilogram.

Construction of appropriately labelled tables

Data is most effectively summarised when presented in the form of a table.

A well-drawn table will include the following features:

- A numerical header such as Table 1.
- A suitable title – including reference to the independent variable being investigated, and the dependent variable which is being measured/observed.
- A heading for each column – ensuring that units are in brackets if quantitative units have been used.

All quantitative observations made are recorded within the table. Data must be recorded to the appropriate and same number of significant figures down each column.

An example of a well-constructed table is shown below in Figure 7.02.

$Na_2S_2O_3$ volume (mL)	H_2O volume (mL)	HCl volume (mL)	Total volume of solution (mL)	Time taken (s)
45	0.0	5.0	50	21.01
40	15	5.0	50	23.10
35	10	5.0	50	25.10
30	15	5.0	50	28.97
25	20	5.0	50	38.44
20	25	5.0	50	47.56
15	30	5.0	50	62.38
10	35	5.0	50	107.2
5.0	40	5.0	50	294.1

Table 1: The time taken for varied concentrations of sodium thiosulfate solution to become opaque after reaction with hydrochloric acid.

Figure 7.02: A sample results table

Tables are constructed for raw and derived data. Depending on the complexity of the constructed table, data trends may be able to be observed from inspection. Some data trends may not be apparent until a graph or further mathematical analysis has been undertaken.

Tables are always labelled and are referred to in the experimental report.

Drawing of graphs

Once raw data has been collected through the course of an experiment, and any derived data calculated, trends in the data, and the relationship between variables, should be explored through the construction of a graph.

Graphs can take a variety of forms, and the chosen form is determined by the type of data collected and the purpose of the data collection. Comparisons are generally illustrated as histograms or bar graphs. For example, when comparing the experimental value to the theoretical value for the enthalpy of solution, the data can be easily visualised using a histogram.

Scientific investigations commonly require the construction of a line graph or scatter plot.

Line graphs

This type of graph is used when one variable (independent) affects another (dependent). They are particularly useful for predicting trends. An example of a line graph is shown in Figure 7.03.

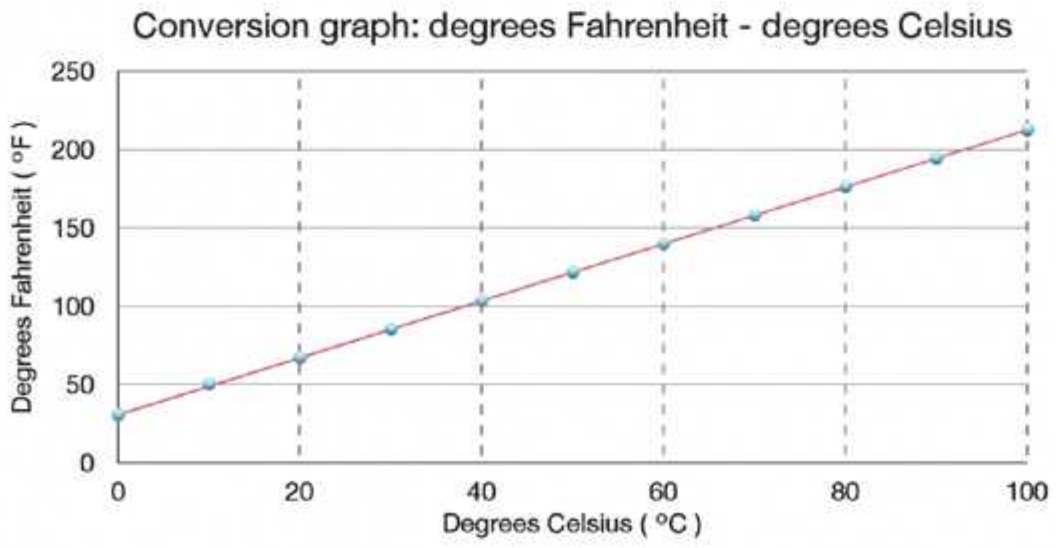


Figure 7.03: An example of a line graph

Scatter plots

Scatter plots are useful for scientific investigations when looking for a pattern in the data. This type of graph can show comparisons. To investigate the relationship of variables, a line of best fit is often applied.

Line of best fit

A **line of best fit** (or 'trendline') is a line that best represents the data on a scatter plot. This line may pass through some of the points, none of the points, or all of the points. A best-fit line is meant to mimic the trend of the data. A scatter plot with a linear line of best fit applied is shown below in Figure 7.04.

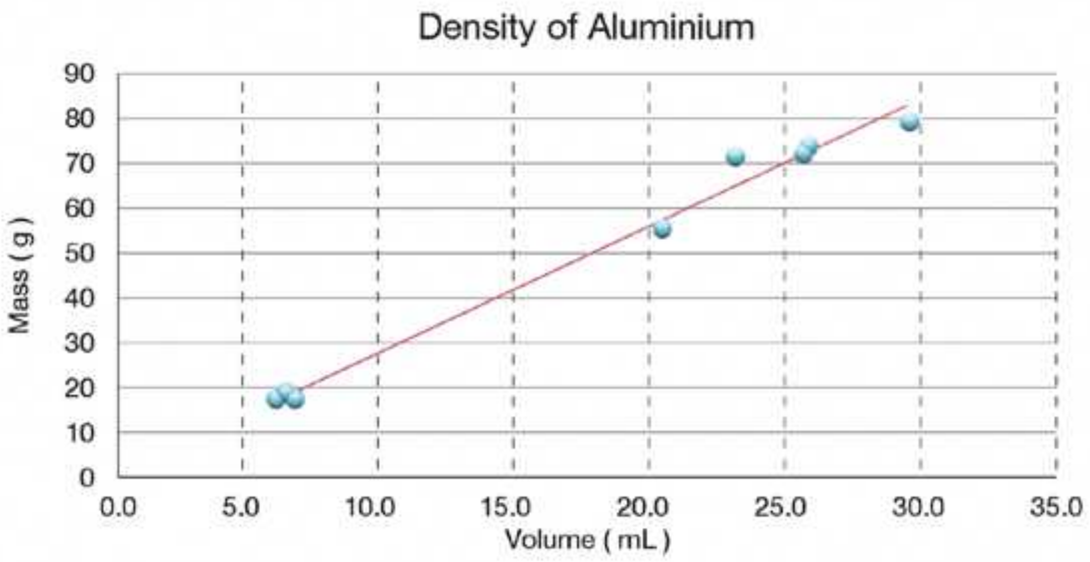


Figure 7.04: An example of a scatter plot

Determining the most appropriate trendline for the collected data is important. Common trendlines are summarised in Table 7.02.

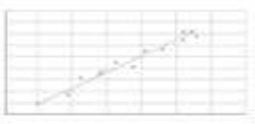
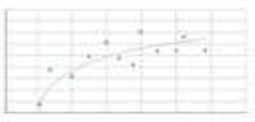
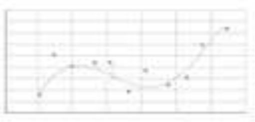
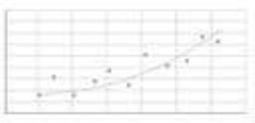
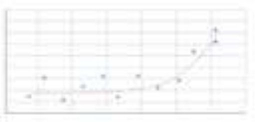
Trend type	Why/when it is used	Example
Linear	A linear trendline is a best-fit straight line that is used with simple linear datasets. A linear trendline usually shows that something is increasing or decreasing at a steady rate.	
Logarithmic	A logarithmic trendline is a curved line of best fit that is most useful when the rate of change in the data increases or decreases quickly and then levels out. A logarithmic trendline can use negative and positive values.	
Polynomial	A polynomial trendline is a curved line that is used when data fluctuates.	
Power	A power trendline is a curved line that is best used with data sets that compare measurements that increase at a specific rate.	
Exponential	An exponential trendline is a curved line that is most useful when data values rise or fall at increasingly higher rates. An exponential trendline cannot be used if the data contains zero or negative values.	

Table 7.02: Common trendlines

Further discussion of the ways in which graphs can be analysed will be discussed in Topic 7.5.

Once the most appropriate type of graph has been determined, the following steps can be undertaken to form a well-constructed graph.

To demonstrate these steps clearly, the below experimental example will be used.

Example

Data from an experiment exploring Boyle's Law is shown in Table 7.03. Boyle's Law states that the pressure of a given mass of an ideal gas is inversely proportional to its volume at a given temperature.

Pressure (kPa)	Volume ($m^3 \times 10^{-6}$)
250	31
180	34
140	44
100	62
85.0	73
70.0	88
60.0	99

Table 7.03: Boyle's Law experimental data

1. **Placement of variables**

   The independent variable is usually placed on the x-axis of a graph, and therefore the dependent variable is placed on the y-axis. The dependent variable is generally positioned on the vertical (y) axis as it is the solution to the equation of the line ($y = mx$).

For Boyle's Law it is the pressure which is varied and the volume which is subsequently measured, therefore pressure will be assigned to the x-axis and volume to the y-axis.

2. **Variable range**

   The numeric range of the data obtained for each variable must be calculated before the scale of the axes can be determined. This determines the range of values that must be represented on each scale.

   Pressure (x-axis) = range of 250 units

   Volume (y-axis) = range of 99 units

3. **Scale**

   The scale is the number value of each square on an axis. It is generally a reflection of the resolution of the measuring instrument used to measure a given variable. To determine the scale two things must be known:

   (a) The number of squares available along each axis of the graph grid.

   (b) The range of the variable to be represented along each axis.

The following 25 × 25 grid space will be used to graph the data, and therefore represents the value required for (a):

To determine the required scale, divide the largest of these two numbers (a and b) by the smallest.

X-axis:

(a) *= 25*

(b) *= 250*

Scale = $\frac{250}{25}$ *= 10 units*

This will give you the value of each square on the axis.

If the division gives a remainder, round off the number so the data will fit on the graph. Spread the graph so that data will cover most of the space.

Y-axis:

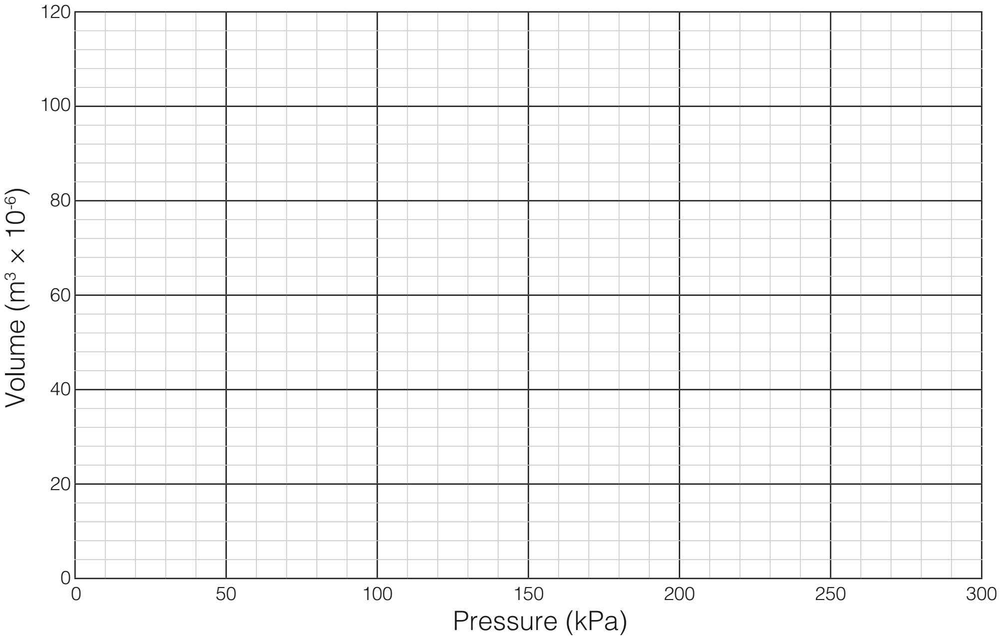
(a) *= 25*

(b) *= 99*

Scale = $\frac{99}{25}$ *= 3.96 units = 4 units*

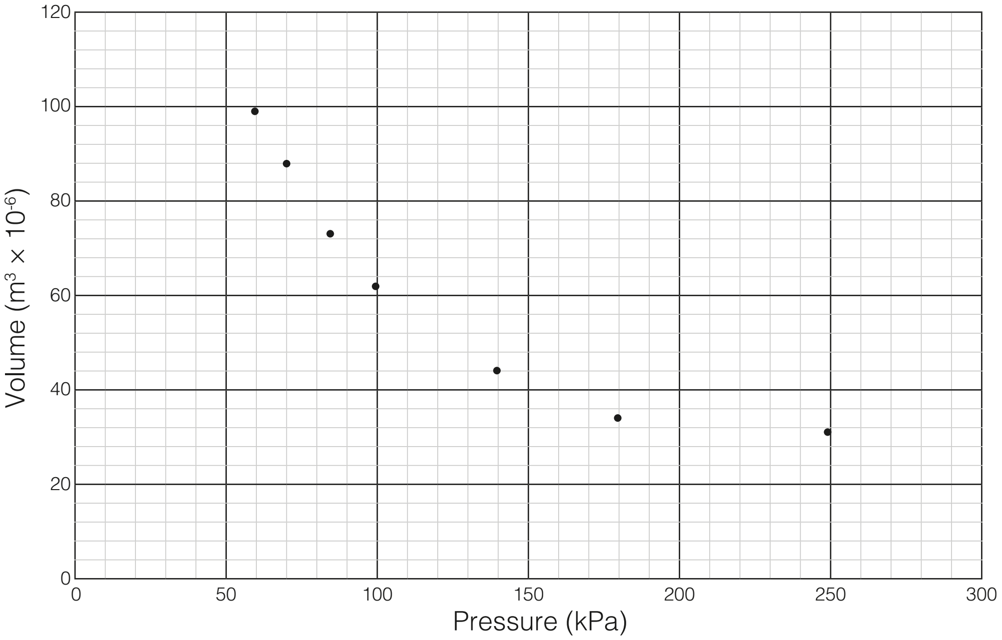
4. Number and label each axis

Using the scale for each axis, write numbers along the axis – increasing from left to right and bottom to top. Remember, the scale does not need to start at zero. Title each axis with the name of the variable, with the units alongside in brackets.



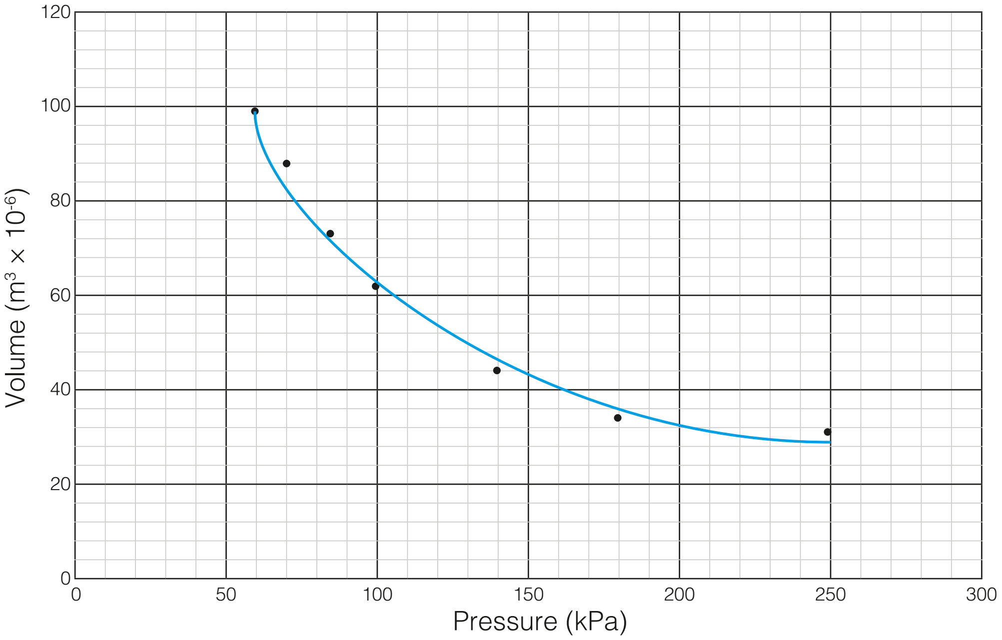
5. Plot the data

Locate each data point with a small dot on the graph.



6. Draw the line of the graph

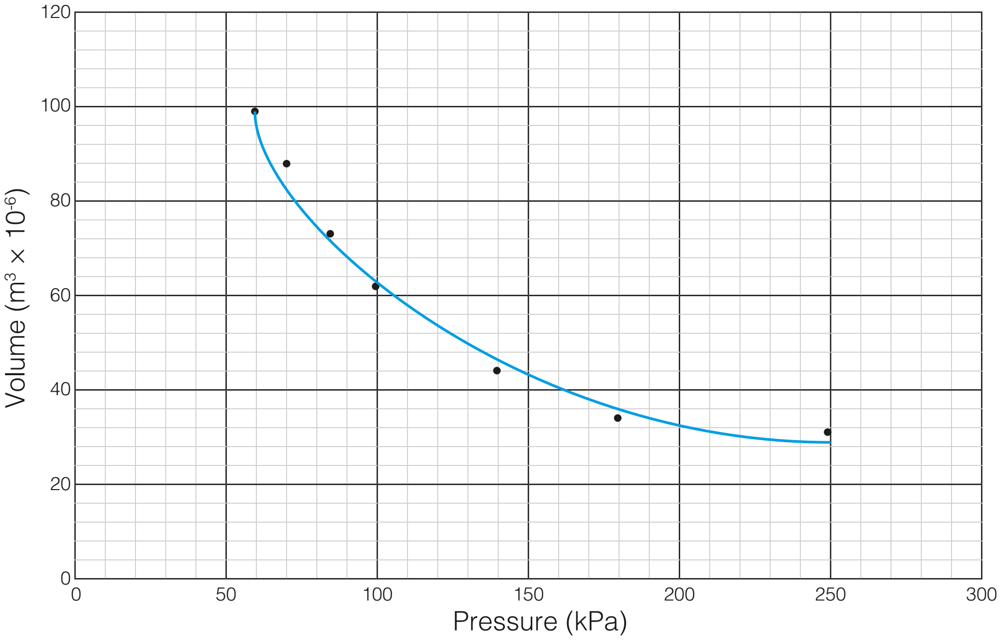
Draw the line or curve that best fits the data points.



7. Title the graph

A graph title should clearly communicate what the graph shows, and reflect the independent and dependent variables.

The inversely proportionate relationship between pressure and volume (Boyle's Law)



Note: If a graph has more than one set of data, a "key" must be included to identify the different lines. Like the title, a graph key should be placed in an open space inside the grid – not in the margin of the paper.

Questions

6. A student decided to boil some water in a beaker. The student used a stopwatch to record the temperature change every minute for ten minutes. A thermometer was then used to measure the temperature change.

   The temperature of the water when he began was 22°C. After the first minute, the temperature was 24°C. It then rose by 6 degrees by the second minute. It rose by 8 degrees every minute for the next three minutes. By the sixth minute, the temperature was 60°C. It then rose 4 degrees for one minute and 6 degrees for the last three minutes.

   (a) Use the information above to construct a table of data from the described experiment.

   (4 marks) **IAE2**

   (b) Define the units used to measure this data as fundamental or derived units. Explain your answer.

   (2 marks) **KA2**

7. The volume of a gas decreases as the temperature of the gas decreases. A sample of gas was collected at 100°C and then cooled. The changes in the volume of the sample are shown below.

Temperature (°C)	100	80	60	40	30	20	10	0	–10	–30
Volume (mL)	317	297	288	278	252	243	236	233	227	202

   (a) Graph the data using the provided space below.

   (6 marks) **IAE2**

8. Ethylene is a plant hormone that causes fruit to mature. The data below concerns the amount of time it takes for fruit to mature from the time of the first application of ethylene by spraying a field of trees.

Amount of ethylene in mL/m^2	**Wine sap apples: Days to maturity**	**Golden apples: Days to maturity**	**Gala apples: Days to maturity**
10	14	14	15
15	12	12	13
20	11	9.0	10
25	10	7.0	9.0
30	8.0	7.0	8.0
35	8.0	7.0	7.0

(a) Graph the data using the provided space below.

(6 marks) **IAE2**

Scientific conventions

Science understanding

Effective scientific communication is clear and concise. *© Copyright SACE 2022.*

Information presented in scientific reports must be clear, concise and well-articulated. Thorough and successful scientific practice does not only emerge from strong conceptual understanding, problem-solving and use of practical equipment but the ability to communicate effectively in multiple forms within a report. The first step to achieving this is to ensure the general scaffold (outlined in Topic 7.2) of a scientific report is followed.

Importantly, scientific communication must be reflective of a good understanding of and ability to implement correct grammar and syntax in extended written prose. Sentences should be complete and grammatically correct, ensuring that (where appropriate) short and clear sentences are used to convey the required information. It is often said that a report should be able to be read and understood by someone with little background knowledge of the subject.

“If you can’t explain it to a six year old, you don’t understand it yourself.” – Albert Einstein

To achieve this, clear communication of ideas and experimental protocols is required. Technical terms need to be used correctly and spelling must be carefully checked. In terms of style, a scientific report needs to be formal and scientifically accurate. It is objective, and devoid of personal pronouns, emotive words and ambiguous expressions.

Science understanding

> Scientific information can be presented using different types of symbols and representations.
>
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Scientific conventions assist chemists to explain concepts, solve problems and make predictions. They are another component of effective scientific communication.

Mathematics in chemistry

Mathematics is used widely in chemistry, as calculations are necessary to explore many important concepts in chemistry. A basic understanding of mathematics is required to successfully interpret the data obtained from an experiment.

Scientific notation

Scientific notation is a way of expressing numbers that are too big or too small to be conveniently written in decimal form. In essence, in scientific notation is written in two parts:

1. The digits are written (with the decimal point placed after the first digit), followed by
2. **× 10 to a power** that places the decimal point where it should be (i.e. it shows how many places to move the decimal point).

This is demonstrated in Figure 7.05 below:

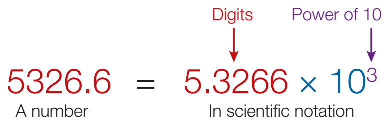


Figure 7.05: Conversion of a natural number to scientific notation

Importantly:

- A power of ten with a positive exponent, such as 10^5, indicates the decimal was moved to the right.
- A power of ten with a negative exponent, such as 10^{-5}, indicates the decimal was moved to the left.

Example

Consider 1.0×10^5

If the decimal place is moved to the right then this results in: 100000.0

If the decimal place is moved to the left then this results in: 0.000010

Ratios

A ratio is a number or pair of numbers representing the size of one quantity relative to another. Ratios are common in chemistry and are encountered in many forms. Fractions and percentages are specific types of ratios. Chemical formulas are an example of specialised ratios.

Part-to-part ratios are commonly written using a colon. Such ratios are common in this discipline (otherwise known as stoichiometry) when representing and using mole ratios to calculate the value of an unknown using a known quantity.

Percentages

Percentages are ratios that can be written as a fraction with the denominator equal to one hundred (representing the whole). This convention is used throughout chemistry. Common examples include mass, volume, and percentage yield.

Parts per million (ppm) and parts per billion (ppb) are terms for ratios smaller than a percent (parts per hundred). For example, the ppm ratio has a denominator equal to one million (10^6) instead of one hundred.

Diagrams

Drawn diagrams are a particularly useful scientific convention for displaying or conveying information effectively. For diagrams to successfully portray or explain a concept they must:

- be scientifically accurate
- have clearly labelled features
- have a clear header such as Figure 1
- be referred to in the text (the same way you refer to graphs and tables)

The terms ‘Figure’ and ‘Table’ are assigned capital letters as they are pronouns.

Diagrams can take a variety of forms including (not exclusively) the types shown in Table 7.04:

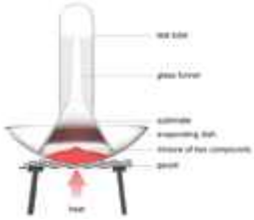
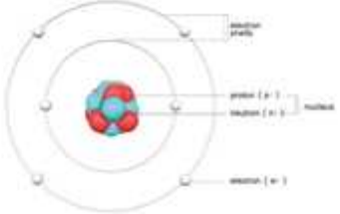
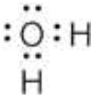
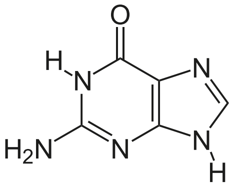
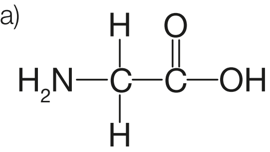
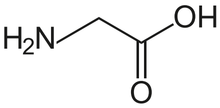
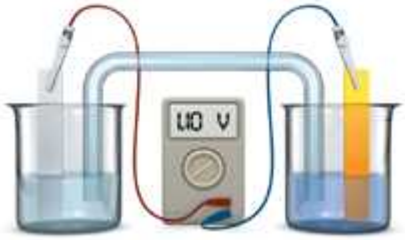
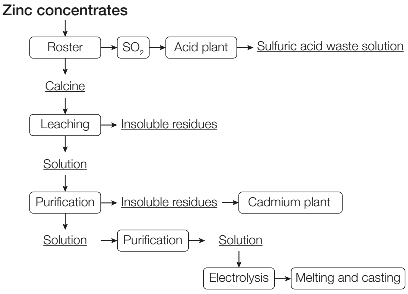
Diagram type	Example
The apparatus involved in an experiment	
Atomic diagrams	
Lewis dot diagrams	
Molecular structures	 guanine (G)
Organic structural formulae: (a) Full (b) Condensed (c) Skeletal	Glycine (a)  $H_2N-\overset{H}{\underset{H}{C}}-\overset{O}{\overset{\Vert}{C}}-OH$ (b) $C_2H_5NO_2$ (c) 
Electrochemical cells	
Flow charts	

Table 7.04: Common examples of scientific diagrams

Diagrams can be represented by:

- hand-drawn figures
- photos
- computer-generated figures
- a combination of the above.

Effective diagrams supplement written explanations, and can help enhance the manner in which ideas are articulated.

Equations

Chemical equations are a specialised convention that displays the changes that occur in a chemical reaction from the reactants to the products. They can have a variety of forms and, in essence, provide a shorthand for displaying what chemical change occurs during a chemical reaction.

Chemical equations are an important component to the way chemists communicate. The most widely used types of equations are displayed below in Table 7.05.

Equation type	Specifics of that equation type	Example
Word	• Reactants and products are written as words separated by a reaction or equilibrium arrow • Ratios of species are not considered	octane + oxygen → carbon dioxide + water
Symbol	• Reactants and products are depicted with chemical formulae, separated by a reaction or equilibrium arrow • States may or may not be included • All species must be balanced • Reaction conditions are shown above or below the reaction arrow	$2Na_3PO_4 + 3MgCl_2 \rightarrow 6NaCl + Mg_3(PO_4)_2$
Ionic	• Species are shown using chemical formulae, however only reacting ions are included (so spectator ions are excluded)	$Zn^{2+} + 2OH^- \rightarrow Zn(OH)_2$
Thermochemical	• All specifics for symbol equations and also inclusive of the following: • All states must be identified and shown • The magnitude and sign of the enthalpy value must be included	$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}\ \Delta H = -802.4\ kJ\ mol^{-1}$

Table 7.05: Forms of chemical equations

The process of writing and balancing chemical equations is covered in Chapter 4.3.

Questions

9. Convert the following values to scientific notation:

(a) 0.0742: ……………

(b) 0.0018: ……………

(c) 0.0004: ……………

(d) –0.58: ……………

(e) 0.5064: ……………

(f) 7006: ……………

(g) –21522: ……………

(h) 215: ……………

(8 marks) **KA4**

Analysis of data

Science understanding

> The analysis of the results of investigations allows them to be interpreted in a meaningful way.
>
>

The tabulated and graphed experimental data is then interpreted by the experimenter or by a third party. Communication of this analysis becomes an important part of the discussion within the report and is imperative if conclusions are to be drawn from the investigation.

Identification and discussion of trends, patterns and relationships

If graphs have been constructed from the collected data, analysis of the trends and patterns presented on the graph is important for describing the relationships shown in the data.

A trend on a line graph suggests more than just a connection between the two variables. A trend implies that the manipulation of (change to) one variable was responsible for the changes in the other variable.

Trends in the data can be identified through an analysis of the trendline on the graph. In Table 7.06 below, common examples of line graph slopes and related interpretations are shown.

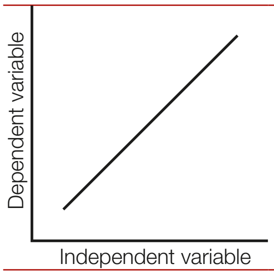
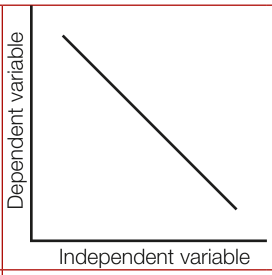
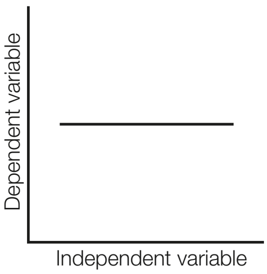
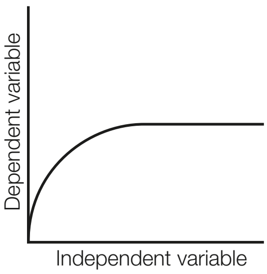
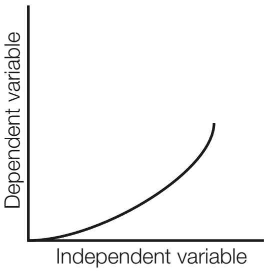
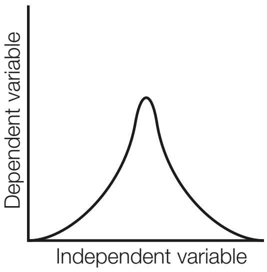
	**Slope:** Constantly rising slope **Interpretation:** Independent variable causes dependent variable to increase regularly		**Slope:** Constantly falling slope **Interpretation:** Independent variable causes dependent variable to decrease regularly
	**Slope:** Constantly level slope **Interpretation:** Independent variable causes no change to dependent variable		**Slope:** Steeply rising slope, gradually lessening to a level line **Interpretation:** Independent variable causes dependent variable to increase initially but the effect decreases until it has no effect
	**Slope:** Slope rises steeper and steeper **Interpretation:** Independent variable causes dependent variable to increase exponentially		**Slope:** Rising slope which peaks to a maximum and then decreases **Interpretation:** Independent variable causes dependent variable to increase to a maximum and then to decrease (i.e. only effect over a narrow range)

Table 7.06: Interpreting slopes of line graphs

Generally, it is said that:

- **Straight lines (linear)** indicate a direct relationship between two variables. Specifically the independent variable is directly proportional to the dependent variable.
- **Curved lines** indicate a changing relationship between the two variables. Every time the Independent Variable changes by a certain amount, the Dependent Variable changes by a differing amount.

The closer the points are to the line of best fit the stronger the correlation is between the variables. In other words, the greater the degree of **scatter**, the greater the influence of **random error** on the experimental data.

Interpolation and extrapolation

Extrapolation and interpolation are both used to estimate hypothetical values for a variable based on other observations. They are useful in predicting values from established trends. There are a variety of interpolation and extrapolation methods based on the overall trend that is observed in the data.

To determine the difference between extrapolation and interpolation the prefixes *extra-* and *intra-* need to be examined. The prefix *extra-* means 'outside' or 'in addition to'. The prefix *intra-* means 'in between' or 'among'.

Extrapolation is the process of predicting values beyond the original observation range. The further away the predicted value is from the plotted values, the less reliable the prediction becomes.

Interpolation is a method of constructing new data points within the range of a discrete set of known data points.

This concept is summarised in Figure 7.06 below.

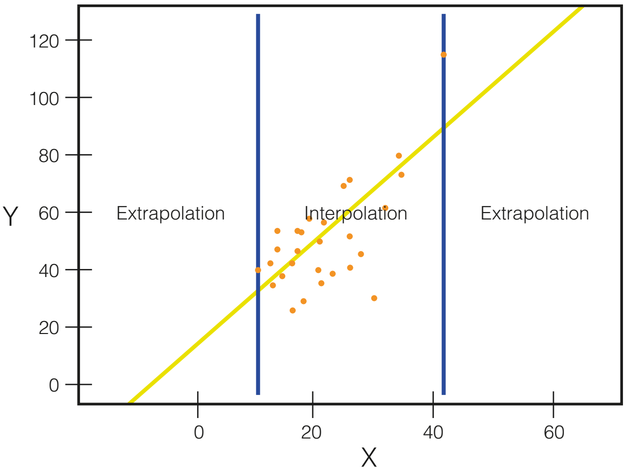


Figure 7.06: A graph representing interpolation between the dataset, and extrapolation extending past the dataset.

Selection and use of evidence and scientific understanding to make and justify conclusions

The analysis of scientific data allows conclusions to be formed. However, it is important that additional evidence supports such conclusions. Research regarding the underlying principles of the investigation, as well as the collection and analysis of similar published investigation reports, is important in providing this. Having additional evidence that supports a chemist's formulated conclusion adds validity and strength to those conclusions presented within that chemist's report. An important skill, therefore, is the ability to critically and effectively select sources of evidence.

Evidence used within a scientific report should be:

- **Suitable** – sources chosen should have a high degree of relevance to the topic, and be suitable to the level of chemistry being presented in the report.
- **Credible** – the authors of work presented should be reviewed to ensure that it is scientifically reliable. Credibility is enhanced by an author's association with reputable tertiary or research-based institutes. Credibility can also be verified through the work's own referencing of ideas and conclusions presented. Credibility can be diminished when no author details are provided, and therefore the origin of the ideas cannot be verified.
- **Accurate** – ideas and conclusions within the source should be chemically correct. This often involves reviewing additional material that supports the concepts explained within a source.
- **Unbiased** – bias within scientific research can be present, and therefore evaluating a source to ensure that it presents an unbiased perspective is important. Where applicable, sources should present a balanced view regarding their conclusions. Bias can be introduced when the origin of a source has something to gain from presenting a particular conclusion. This relates to the ethics of investigations, in that, even if a source is biased, it should never contain information that is falsified or data that is skewed in order to provide a particular, favoured conclusion.

Once sources have been selected, it is important to present evidence from these sources in an appropriate manner. This involves paraphrasing the ideas presented and ensuring that the authors of the original work are acknowledged via an appropriate in-text referencing convention (usually Harvard style or, less commonly, footnotes). In addition to this, all work referred to should be included in a correctly structured reference list.

The process of critical selection and accurate acknowledgement of previous work is not limited to investigation reports. The same process is employed in essays, investigations of human endeavour and other research-based tasks.

Questions

10. Dinitrogen pentoxide decomposes into nitrogen and oxygen via the following reaction:

$$2N_2O_{5(g)} \rightarrow 4NO_{2(g)} + O_{2(g)}$$

To complete this investigation, N_2O_5 was introduced in to a reaction vessel at an initial concentration of 0.100 mol L^{-1}. With a temperature maintained at 65°C, the concentration of nitrogen dioxide was measured at 100 second intervals and produced the following data.

Time (seconds)	**Concentration of NO_2 (mol L^{-1})**
0.00	0.000
100	0.082
200	0.130
300	0.158
400	0.174
500	0.185
600	0.191

(a) Using the above data, graph the formation of NO_2 over the first 600 seconds of the reaction.

(5 marks) **IAE2**

(b) Use the graph to predict the concentration of NO_2 at 425 seconds.

(1 mark) **IAE3**

(c) Use the graph to predict the concentration of NO_2 at 650 seconds.

(1 mark) **IAE3**

(d) Explain how the data predicted in (c) and (d) represents interpolation and extrapolation.

(2 marks) **KA2**

11. Oxygen gas (O_2) can be generated by the reaction of hydrogen peroxide (H_2O_2) with manganese dioxide (MnO_2).

$$2H_2O_2 + MnO_2 \rightarrow 2H_2O + 2O_2$$

A student set up nine test tubes and in each one placed differing masses of MnO_2. An equal amount of H_2O_2 was added to each test tube, and the volume of O_2 produced was measured each minute, for five minutes. The data obtained is summarised below:

Test Tube	**MnO_2 (g)**	**Volume of O_2 produced (mL)**				
		1 minute	**2 minutes**	**3 minutes**	**4 minutes**	**5 minutes**
1	0.1	1.4	2.6	3.5	4.2	5.1
2	0.2	2.8	4.6	5.8	7.1	7.6
3	0.3	4.9	7.2	8.8	10	11
4	0.5	5.9	8.5	10	12	13
5	1.0	8.5	12	14	16	17
6	1.5	11	15	18	20	22
7	2.0	12	17	20	23	25
8	2.5	13	19	22	25	27
9	3.0	16	22	25	28	30

(a) Propose a testable hypothesis for this experiment.

(2 marks) **KA1**

(b) State the volume of O_2 produced in:

(i) Test tube 3 produce between the second and fourth minute.

(1 mark) **IAE3**

(ii) Test tube 5 during the first two minutes.

(1 mark) **IAE3**

(iii) All nine test tubes at the end of the experiment (in L).

(2 marks) **IAE3**

(c) Graph the data for test tubes 2, 4 and 6 using the space below.

(5 marks) **IAE2**

(d) Use the graph to predict the volume of O_2 that will be produced after:

(i) 6 minutes in test tube 2.

.... *(1 mark)* **IAE3**

(ii) 3.5 minutes in test tube 4.

.... *(1 mark)* **IAE3**

(e) Identify the type of trendline represented on the graph constructed in part (c).

.... *(1 mark)* **KA1**

(f) Describe what this type of trendline implies about the relationship between the independent and dependant variable.

....

.... *(1 mark)* **KA1**

(g) State a suitable conclusion for the trends observed on the graph.

....

.... *(1 mark)* **KA1**

Critical evaluation of procedures and outcomes of investigations

Science understanding

Critical evaluation of procedures and outcomes can determine the meaningfulness of conclusions.

Following the analysis of experimental results, evaluation of the experimental procedure is important for identifying sources of certainty and, once again, ensuring conclusions formed from an investigation are valid. Components of such evaluations are commonly communicated within the discussion of the report. Factors that should be considered are explained below.

Mistakes

It is important to distinguish the difference between a mistake and error. A measurement can be made without mistake, but no measurement is without some degree of error. All that can be done is to minimise the error or eliminate the source of the error.

Every measurement is affected by random and or systematic errors. Increasing the number of samples minimises the effects of random errors and increases the reliability of the data. Systematic errors can be identified and results verified by repeating the experiment, using an alternative source of equipment and materials.

Random errors

Random errors are chance variations between successive measurements over which the experimenter has little or no control. The magnitude of random errors is linked to the limitations of the measuring instruments. They follow no pattern and can include:

- Parallax error, which is induced by variations in an individual's eye level when determining the measurement of a solution
- Reaction time error, where, if an individual is responsible for initiating a timer or similar device, the reaction time of that individual will have an effect on the recorded measurements
- Error in estimating the position of a needle between divisions on a scale
- Changes in the quantity being measured due to fluctuations in temperature
- Other sources of random error, which are often connected to specificities of a particular type of investigation.

Taking the average of several repeated measurements can reduce the effect of random error on the dataset as a whole, and thus increases the reliability of the results. If graphed, the line of best fit reduces the effect of random errors. And generally, the greater the degree of scatter, the greater the influence of random error on the collected data.

In terms of communicating random errors within a scientific report, the following three-sentence structure can be useful:

1. State the random error.
2. Explain the effect of the error on the collected results of the investigation.
3. Describe how the effect of the error can be minimised.

Example

Individual's reaction time

This experiment involved the use of a digital stopwatch. This instrument requires an individual to initiate the measurement and, therefore, the reaction time of the experimenter is a possible random error. If the person initiating the stopwatch has a delayed reaction time, then the recorded measurement of time will be increased, decreasing the precision of the collected data. The effect of such an error can be minimised by ensuring the same person initiates the stopwatch throughout the trials. The experimenters should also complete multiple trials so an average can be calculated.

Systematic errors

Systematic errors are recurrent errors inherent in the apparatus or experimental method. They are linked to the accuracy of measurements. A measurement is accurate to the extent that systematic errors have been identified and eliminated in subsequent trials. Systematic errors in chemical investigations can come from:

- A measuring instrument that is incorrectly calibrated.
- A piece of apparatus that has an inherent fault.
- Contaminated standards which are used in quantitative analysis.
- The use of non-homogenous solution.

Systematic errors cannot be minimised, and can only be eliminated at the source. Most commonly this will involve repeating the experiment with new apparatus and materials.

In terms of communicating systematic errors within a scientific report, the following three-sentence structure can be useful:

1. State the systematic error.
2. Explain the effect of the error on the collected results of the investigation.
3. Describe how the effect of the error could be identified should the experiment be repeated.

Example

Contamination of stock solution

This experiment required the use of a standardised solution. If the solution was not correctly inverted before a sample was collected, then the solution would not be homogeneous, resulting in an inconsistent concentration. The result of this may be that the collected sample may be of lower concentration than the stated concentration, lowering the accuracy of the collected data. To eliminate this error, if the experiment is repeated, the bottle must be inverted prior to the sample being collected.

Precision

Precision is the ability of a measurement to be consistently reproduced. The precision of the practical is dependent on the extent to which random errors can be minimised.

The degree of precision of a series of measurements is dependent on the:

- extent to which random errors have been minimised.
- skill of the person taking the measurements.

Precision is indicated by the degree of uniformity of the measurements.

Resolution

Resolution is the term used to describe the minimum value to which a piece of data can be resolved using a given measuring instrument.

Too often precision and resolution are used interchangeably, however, they are very different terms.

Example

Distinguishing between precision and resolution – a stopwatch

Figure 7.07 shows two stopwatches, one digital and the other analog, both operated by an experimenter.

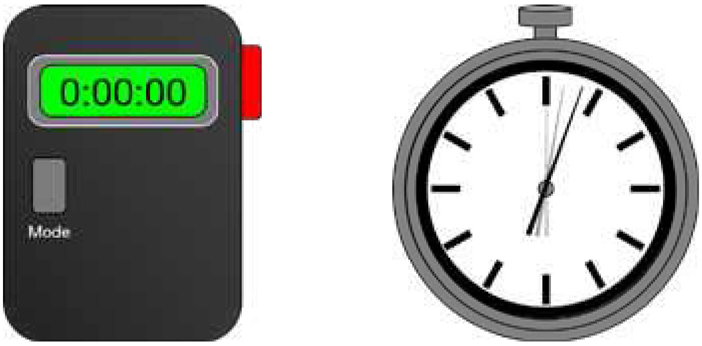


Figure 7.07: Digital and analog stopwatches

From this figure both the resolution and precision of the devices can be evaluated.

Resolution:

- The analog stopwatch has to be viewed on its dial. Upon close view the best a good eye can do is resolve a reading to $\frac{1}{10}$th second, which is, therefore, the resolution of the stopwatch.
- The digital stopwatch has two digits beyond the seconds, so it subdivides time in hundredths of a second. Its resolution is therefore $\frac{1}{100}$th of a second.

Comparing the two devices, there is a substantial difference between the stopwatches in resolution: a power of ten, from $\frac{1}{10}$th to $\frac{1}{100}$th of a second. *Therefore, the digital device has the greatest resolution.*

Precision:

As previously described, precision means *reliable*, *repeatable* measurements. Experiments have shown that a human takes about $\frac{1}{10}$th of a second to react to a stimulus before starting a stopwatch. Therefore,

- the analog stopwatch has a precision of about $\frac{1}{10}$th second. Both the resolution and the stimulus-response time of the human are $\frac{1}{10}$th second.
- The digital stopwatch also has a precision of $\frac{1}{10}$th of a second, however, maintains its resolution of $\frac{1}{100}$th of a second.

Independent of the type of device used, the measurement is *repeatable* to $\frac{1}{10}$th of a second, and *therefore there is no difference in the precision of either device.*

Example

Glassware and resolution – volumetric flask, volumetric pipette, and burette

During titrations (and other experiments that require a high degree of accuracy), selection of the most appropriate measuring equipment is imperative. Some glassware is termed 'volumetric', which refers to glassware that only measures a predefined volume of solution. These are most commonly volumetric flasks, pipettes, and burettes, and are characterised by high accuracy and thus excellent repeatability of measurements. The features of these types of glassware can be explored to reinforce the idea of resolution.

Meniscus

Importantly when measuring the volume of a solution in any measuring apparatus, the bottom of the meniscus must be precisely on a calibration mark.

Note: The formation of a meniscus is due to the interaction of water and glass. Specifically, when water is confined in a tube, its surface has a concave shape because water wets the surface and creeps up the side. If the liquid molecules are attracted more strongly by the glass wall (adhesion) than by their kind (cohesion), the meniscus is curved downwards (concave) and the edge of the liquid surface is slightly raised. The curvature of a meniscus is related to the surface tension of the liquid and inversely related to the diameter of the tubing in which it is formed.

This is shown in Figure 7.08.

In addition to this, to avoid the introduction of parallax error, the solution measurement must be taken at eye level. A clear meniscus indicating a volume of 1.60 mL is shown in Figure 7.09.

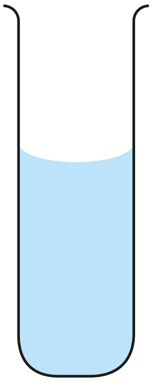


Figure 7.08: Formation of a meniscus

Figure 7.09: Demonstration of a meniscus placement of a calibration mark

Volumetric flask

Volumetric flasks are commonly used as the vessel in which a standard solution is prepared. This process was explained in Topic 4.3. Volumetric flasks are calibrated to predetermined volumes, which allows for accurate preparation of solutions.

A volumetric flask, similar to the one in Figure 7.10 is manufactured to 250 mL ±0.2 mL. The resolution of the flask is equal to that of its volume, and therefore is 250 mL.

Figure 7.10: Volumetric flask

Volumetric pipette

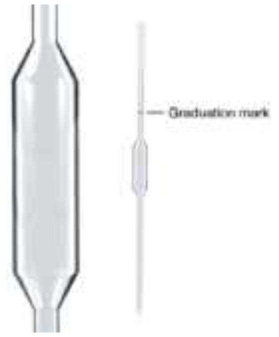


Figure 7.11:
25 mL volumetric pipette

Pipettes are volumetric instruments used for measuring volumes of liquid that are calibrated ‘to deliver’. This calibrated volume of solution is commonly referred to as the **aliquot**. During the manufacturing process, they are individually volumetrically calibrated and provided with one or more calibration marks. An example of a bulb pipette is shown in Figure 7.11.

Like volumetric flasks, bulb pipettes have a resolution equal to their volume.

Burette

A burette is a device commonly used in analytical chemistry such as titrations. They are capable of dispensing small and precise volumes of solutions via a stopcock. A burette consists of a graduated glass tube, usually with a resolution of 0.1 mL. A 50 mL burette is depicted in Figure 7.12.

Figure 7.12: 50 mL burette

Science as a human endeavour

Titrations

The appropriate selection and correct use of the above glassware is imperative in titrations. Titration refers to a chemical technique involving the slow addition of one solution of a known concentration (the titrant) to a known volume of another solution of unknown concentration (the aliquot). Titrations are classified based on the type of reaction that takes place between the titrant and the aliquot. The most common titrations involve acid–base and redox reactions.

The origin of titrations dates back to late 18th century French chemistry. The first burette was developed by Francois Antoine Henri Descroizilles in 1791, and its design similar to that of a simple graduated cylinder. In 1824, Joseph Louis Gay-Lussac improved the design to include a side arm. He was also responsible for introducing the terms ‘burette’ and ‘pipette’ into common chemical terminology. The methodology of titration underwent significant improvement due to the work of Karl Friedrich Mohr in around 1855; he designed the burette to include a clamp and tip at the bottom. He wrote the first text on the technique, *Lehrbuch der Chemisch-analytischen Titrirmethode* (Textbook of analytical-chemical titration methods).

Accuracy

The accuracy of an experimental value indicates how close the collected data is to the theoretical or literature value, and depends on the extent to which systematic errors are eliminated.

The degree of accuracy of a series of measurements is:

- Determined by the degree of conformity of the measurements with the accepted or true value of the measurements
- Increased if the measuring instrument is calibrated or a standard solution is standardised before use

The relationship between precision and accuracy is depicted in Figure 7.13.

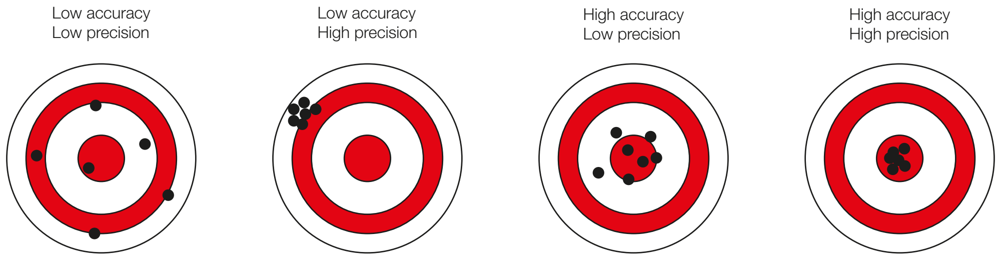


Figure 7.13: Precision/accuracy relationship

Example

Rinsing glassware

The manner in which laboratory techniques are carried out can influence an experiment's precision and accuracy. The way in which a meniscus is read can affect the precision of the measurement taken, and the techniques employed to rinse glassware can greatly influence the overall accuracy of a practical. Therefore, an understanding of rinsing techniques is imperative. To highlight this, the rinsing required for volumetric pipettes and burettes will be outlined below.

> *Both volumetric pipettes and burettes must first be rinsed with distilled water. This process eliminates the presence of any contaminants from previous experiments or dishwasher rinses (most commonly alkaline solutions).*
>
> *After this initial rinse, both pieces of apparatus must be rinsed again with the solution that they are to contain for the experiment.*

A volumetric pipette is rinsed via the following steps:

1. Attach a bulb filler to the top of the pipette, and depress the bulb as shown in Figure 7.14.

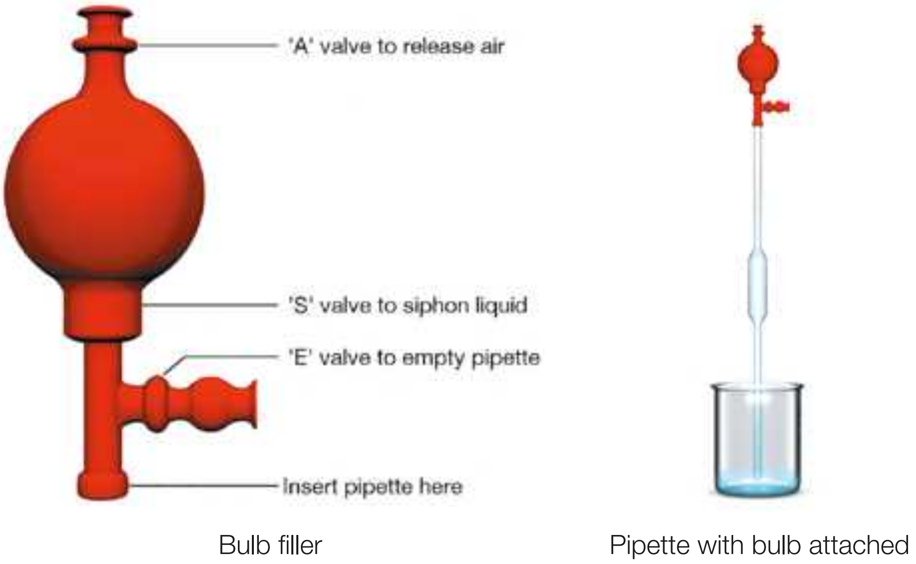


Figure 7.14: Equipment required to fill a volumetric pipette

2. Place the pipette into a beaker of the solution that will be used for the rinse. The bulb filler is then used to siphon the solution into the pipette *past* the calibration line.
3. Vertically evacuate the solution into a waste container by removing the bulb filler.

A burette is rinsed via the following steps:

1. Fill the burette one third of the way with the solution that will be used for the rinse (ensure the stopcock is horizontal and thus closed).
2. Use the horizontal swirl technique. This involves lowering the burette to a horizontal position and rotating the burette on its axis. This ensures that the solution rinses all of the internal surfaces of the burette.
3. Eliminate a portion of the rinsing liquid via the open end of the burette.
4. Use the stopcock to evacuate the remaining liquid with the burette in the vertical position over a waste container.

Improvements

In addition to evaluating possible sources of errors, and discussing the precision and accuracy of the collected data, it is useful to consider and communicate possible improvements to the designed experimental method. Improvements should be appropriate for producing data that will be of equal or greater value when forming a conclusion. Improvements may involve simple changes to the quantities involved or the reaction conditions applied, or more significant changes to the method itself or the variable being investigated.

Reliability and validity

A well-designed investigation should be reliable. **Reliability** refers to the fact that the data should be able to be reproduced if the same method is repeated. Therefore, the reliability of data can be increased by repeating the procedure. **Validity** refers to the credibility of the investigation. It is the extent to which the processes and resultant data measured what was intended or hypothesised.

For conclusions of investigations to be recognised and supported, they should be both reliable and valid.

Questions

12. A student was concerned that the thermometer used in an experiment was not giving satisfactory results. The student tested her thermometer and two other thermometers in a solution at a temperature of exactly 50.00°C and obtained the following readings:

Thermometer used: 49.90°C

Thermometer A: 50.35°C

Thermometer B: 50.20°C

(a) Identify the thermometer with the greatest resolution.

.... *(1 mark)* **KA1**

(b) Identify the most accurate thermometer.

.... *(1 mark)* **KA1**

13. Three students (A, B and C) each weighed the same carboxylic acid sample using three different balances. The correct mass of the carboxylic acid was known to be 22.70 g. the results are shown in the table below:

Weighing	Brad (g)	Shane (g)	Natalie (g)
1	26.1	22.70	23.10
2	23.2	22.75	23.15
3	24.7	22.75	23.15

(a) Define the term ‘precision’.

....

....

.... *(2 marks)* **KA1**

(b) Comment on the scatter of results of students Brad, Shane and Natalie.

....

....

.... *(2 marks)* **IAE3**

(c) Compare the accuracy of the results of students Brad and Shane.

....

....

.... *(2 marks)* **IAE3**

(d) Comment on the resolution of the results of students Brad, Shane and Natalie.

(2 marks) **KA2**

14. A student was concerned that the electronic balances she used in the laboratory were not giving satisfactory results. She then tested two electronic balances in the laboratory with a calibrated weight of exactly 20.0 g, taking four readings each time. She recorded the following results:

Balance 1:	20.001 g	20.002 g	20.001 g	20.001 g
Balance 2:	20.100 g	19.980 g	20.000 g	19.990 g

(a) Identify the most precise balance.

(1 mark) **KA1**

(b) Identify the most accurate balance.

(1 mark) **KA1**

(c) Identify the balance which needs calibrating as it is producing a systematic error.

(1 mark) **KA1**

(d) Comment on the resolution of the balances used by the student.

(2 marks) **KA2**

15. The density of water at 20°C is 0.99823 g mL^{-1}. A student attempts to weigh 1 mL of water at 20°C three times and obtains readings of 0.9420, 0.9846 and 0.9844 g. State the type of error these measurements represent and justify your response.

(2 marks) **IAE4**

16. An Australian Gold Koala coin from the Perth Mint has a minimum gross weight of 3.108 g and is 99.99% pure gold. This coin is weighed three times on four different sets of scales (including the scale housed at the Mint) and the following results are obtained.

	Mint Scale (g)	**Scale 1 (g)**	**Scale 2 (g)**	**Scale 3 (g)**
Trial 1	3.1082	3.11	3.216	3.1
Trial 2	3.1083	3.10	3.215	2.7
Trial 3	3.1082	3.12	3.215	3.3
Significant figures				

(a) Complete the row in the table above, indicating the number of significant figures each column of data is expressed to.

(2 marks) **KA4**

(b) Using examples from the data above, explain the term accuracy.

(2 marks) **KA2**

(c) Using examples from the data above, explain the term precision.

(2 marks) **KA2**

Summary Test 7: Science inquiry skills

1. An experiment was carried out to observe the reaction rate between hydrochloric acid and marble chips (calcium carbonate) at different temperatures.

   The temperature of each flask and the volume of gas produced in one minute in each flask were recorded. The table below shows the results of the experiment.

Flask	Temperature (°C)	Volume of carbon dioxide (mL)
1	24.1	5.00
2	37.5	11.2
3	45.7	21.4
4	61.2	46.4
5	69.8	82.8

(a) State one hypothesis that could be tested in the experiment.

(2 marks) **KA1**

(b) State whether or not the results support the hypothesis written for part (a). Explain your answer.

(2 marks) **KA2**

(c) Identify two factors that should be held constant in the five flasks.

(2 marks) **KA1**

(d) State and explain the effect of one random error that may occur during this experiment.

(2 marks) **IAE4**

(e) Indicate one way the experimenter can reduce the effect of random errors.

(1 mark) **IAE4**

(f) State and explain the effect of one systematic error that may occur during this experiment.

(2 marks) **IAE4**

(g) Indicate one way the experimenter could eliminate the described systematic error in (f) if the experiment was to be repeated.

(1 mark) **IAE4**

(h) Draw a graph for the tabulated data on the grid provided below:

(5 marks) **IAE2**

(i) Use your graph to predict the volume of gas produced at:

(i) 30 °C. *(1 mark)* **IAE3**

(ii) 50 °C. *(1 mark)* **IAE3**

Total: 19 marks

2. To calculate the experimental value for the ΔH neutralisation of nitric acid and sodium hydroxide, the following procedure was performed.
   1. 30.0 mL of 0.150 mol L^{-1} sodium hydroxide was placed in a 200 mL beaker and its temperature was recorded.
   2. 30.0 mL of 0.150 mol L^{-1} nitric acid was placed in a 100 mL beaker and its temperature was recorded.
   3. The nitric acid was rapidly added to the vigorously stirred sodium hydroxide and the maximum temperature was recorded (Experiment 1).
   4. The experiment was repeated but the sodium hydroxide was held in a 200 mL polystyrene cup (Experiment 2).
   5. The third experiment used the same procedure but the sodium hydroxide was held in a 200 mL insulated flask (Experiment 3).

Experimental data

Experiment	Initial temperature (°C) sodium hydroxide	Initial temperature (°C) nitric acid	Final temperature (°C) of reaction mixture
1	14.0	14.0	14.3
2	14.0	15.0	15.5
3	14.0	14.0	15.0

(a) State the aim of the experiment.

....

....

.... *(2 marks)* **KA2**

(b) State a suitable hypothesis for the experiment.

....

....

.... *(2 marks)* **KA2**

(c) State the dependent variable in the experiment.

.... *(1 mark)* **KA1**

(d) State the independent variable in the experiment.

.... *(1 mark)* **KA1**

(e) State two controlled factors in the experiment.

....

....

.... *(2 marks)* **KA1**

(f) Provide a materials and apparatus list for this experiment.

(2 marks) **KA4**

(g) Describe one aspect of safety that would need to be considered in this practical. Include how to minimise any presented risk.

....

....

.... *(2 marks)* **KA2**

(h) State and explain the effect of a possible systematic error in this experiment.

....

....

.... *(2 marks)* **IAE3**

(i) If the accurate temperature change was known to be 1.2°C, state which experiment was most accurate and explain your choice.

....

....

.... *(2 marks)* **KA2**

Total: 16 marks

3. Several students performed a class experiment to look at the best conditions for preparing an organic compound. The experiment involved refluxing (gentle controlled heating over a extended period of time) three chemicals for 30 minutes and then isolating the organic product. The class results are shown below:

Student	**Amount of reactants refluxed**			**Amount of organic product obtained**
	Alcohol (g)	**Organic acid (g)**	**Inorganic acid (g)**	**Ester (g)**
A	10	30.0	2.00	25
B	10	30.0	2.50	27
C	10	30.0	3.00	28
D	10	30.0	3.50	32
E	10	30.0	4.00	31
F	10	30.0	4.50	30

(a) State which set of weights of reactant chemicals suggests it was measured with a mass balance of lower resolution.

(1 mark) **KA1**

(b) State the aim of the experiment.

(2 marks) **KA2**

(c) State a hypothesis of the experiment.

(2 marks) **KA2**

(d) State the independent variable of the experiment.

(1 mark) **KA1**

(e) State the dependent variable of the experiment.

(1 mark) **KA1**

(f) Describe a source of systematic error that could be present in this experiment. Explain the subsequent effect of this error.

(2 marks) **IAE3**

(g) Describe two possible improvements that could be made to the experiment.

(2 marks) **IAE4**

(h) State a conclusion which can be made from this experiment.

(2 marks) **IAE3**

Total: 13 marks

4. A titration was carried out to confirm the concentration of commercial hydrogen peroxide, using standardised potassium permanganate solution. This analysis required the selection and use of specialised glassware, specifically a volumetric flask, volumetric pipette and burette.

   (a) Describe why a volumetric pipette is used to measure the required aliquot of solution instead of a measuring cylinder.

   (2 marks) **KA2**

   (b) Explain the steps required in cleaning a burette before it is filled and used in the titration. Ensure you describe why each step is imperative for maintaining the accuracy of the experiment.

   (4 marks) **KA4**

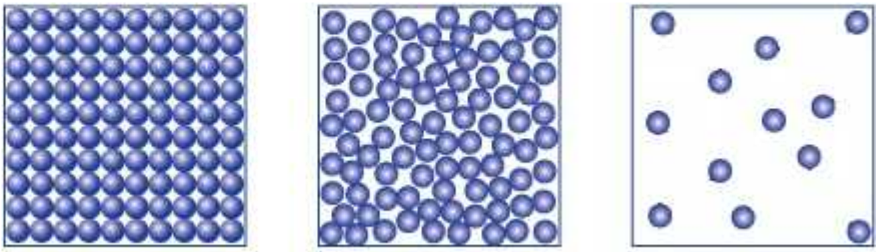
   (c) The student carrying out the experiment rinses the pipette with only distilled water. Describe the type of mistake that this introduces and the possible effect this will have on the results of the titration.

   (2 marks) **IAE3**

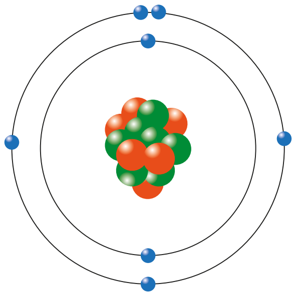
Total: 12 marks

Solutions

Topic 1: Materials and their Atoms

<table>
<tr><th>Question</th><th>Part</th><th>Author's response</th><th>Marks</th></tr>
<tr><td rowspan="3">1</td><td>(a)</td><td>Solid Liquid Gas</td><td>3</td></tr>
<tr><td>(b)</td><td>The particles that comprise liquids can move relative to one another, so the liquid can change shape.
The particles that comprise liquids cannot break away from the other particles, so the volume of the liquid cannot change.</td><td>1
1</td></tr>
<tr><td>(d)</td><td>The boiling point of a material is the temperature at which the material transitions from liquid phase to gas phase.</td><td>1</td></tr>
<tr><td>2</td><td></td><td>
<table>
<tr><th>Material</th><th>Use</th><th>Property/Properties</th></tr>
<tr><td>aluminium</td><td>cookware</td><td>high thermal conduction
high melting point</td></tr>
<tr><td>plastic</td><td>cookware handles</td><td>low thermal conduction
relatively high melting point</td></tr>
<tr><td>wool</td><td>gloves</td><td>low thermal conduction</td></tr>
<tr><td>copper</td><td>wiring</td><td>high electrical conductivity</td></tr>
<tr><td>nichrome</td><td>heating elements</td><td>can conduct electricity
high melting point</td></tr>
<tr><td>Kevlar</td><td>bulletproof vests</td><td>high strength/impact resistance</td></tr>
<tr><td>gelatin</td><td>medicated gel capsules</td><td>partially soluble in water</td></tr>
</table>
</td><td>1
1
1
1
1
1
1</td></tr>
<tr><td rowspan="3">3</td><td>(a)</td><td>Particles in materials are in constant motion and frequently collide with each other.
These collisions transfer heat energy between particles, resulting in transmission of the heat energy throughout the material.</td><td>1
1</td></tr>
<tr><td>(b)</td><td>The particles in a gas are separated by relatively large distances;
this leads to less frequent collisions between particles and reduced heat transfer.</td><td>1
1</td></tr>
<tr><td>(c)</td><td>Fibreglass contains only very thin glass fibres, but many tiny pockets of air.
The air, not the glass, acts as the insulating material.</td><td>1
1</td></tr>
<tr><td rowspan="2">4</td><td>(a)</td><td>The grains of titanium dioxide in sunscreen are each smaller than 100 nm, which meets the definition of a nanomaterial.</td><td>1</td></tr>
<tr><td>(b)</td><td>They reflect and scatter UV light.</td><td>1</td></tr>
<tr><td rowspan="2">5</td><td>(a)</td><td>Smaller than 100 nm;
the cobalt particles must be smaller than 100 nm to qualify as nanoparticles.</td><td>1
1</td></tr>
<tr><td>(b)</td><td>If stronger aluminium is used, less aluminium overall will be required to make the fuselage strong enough to be safe.
If less metal is used in the fuselages of new planes, the mass of the planes will be decreased, and less fuel will be needed to achieve airspeed.</td><td>1
1</td></tr>
</table>

<table>
<tr><th>Question</th><th>Part</th><th>Author's response</th><th>Marks</th></tr>
<tr><td rowspan="2">6</td><td>(a)</td><td>A 'nanotechnology' is any practical application of nanomaterials.
Nanomaterials have important structural features smaller than 100 nm.</td><td>1
1</td></tr>
<tr><td>(b)</td><td>The incorporation of carbon nanotubes has increased the efficiency of solar panels.
This makes the panels more affordable and/or more effective at generating electricity.
or
Nanostructured sponges can absorb many times their own weight in pollutants such as crude oil.
This allows for rapid clean-up of oil spills, reducing the impact on the environment and economy.
or
Any specific benefit linked to a specific nanotechnology is acceptable.</td><td>1
1</td></tr>
<tr><td>7</td><td></td><td><table>
<tr><td>clear apple juice</td><td>homogeneous mixture</td></tr>
<tr><td>baking soda ($NaHCO_3$)</td><td>compound</td></tr>
<tr><td>air</td><td>homogeneous mixture</td></tr>
<tr><td>helium</td><td>element</td></tr>
<tr><td>beach sand</td><td>heterogeneous mixture</td></tr>
<tr><td>concrete</td><td>heterogeneous mixture</td></tr>
<tr><td>carbon dioxide</td><td>compound</td></tr>
<tr><td>milk</td><td>heterogeneous mixture
<i>(globules of fat and protein are distinct at the microscopic scale)</i></td></tr>
<tr><td>sugar ($C_{12}H_{22}O_{11}$)</td><td>compound</td></tr>
<tr><td>ice cream sundae</td><td>heterogeneous mixture</td></tr>
</table>One mark for each correct answer.</td><td>10</td></tr>
<tr><td rowspan="4">8</td><td>(a)</td><td>The mixture is heterogeneous.
Plus, any one of:
The parts of the mixture are different to one another.
The mixture contains different phases of matter.</td><td>1
1</td></tr>
<tr><td>(b)</td><td>Filtration.
The mixture is passed through a filter, which retains the solid lead iodide particles while permitting the flow of water and dissolved potassium iodide ions.</td><td>1
1</td></tr>
<tr><td>(c)</td><td>The mixture is heated in an evaporating dish.
Water has a much lower boiling point than potassium iodide, so it will evaporate, and the potassium iodide will be retained in the dish.</td><td>1
1</td></tr>
<tr><td>(d)</td><td>Distillation (or simple distillation)</td><td>1</td></tr>
<tr><td rowspan="3">9</td><td>(a)</td><td>Boiling point</td><td>1</td></tr>
<tr><td>(b)</td><td>The mixture is heated in a flask attached to a condenser and another flask.
Methanol has a lower boiling point than ethanol, so it will boil, condense in the condenser, and collect first in the receiving flask.
The ethanol will boil at a higher temperature and can be collected in a different receiving flask.</td><td>1
1
1</td></tr>
<tr><td>(c)</td><td>Pure substances are elements, or;
compounds formed from atoms that are chemically bonded in a fixed ratio.</td><td>1
1</td></tr>
</table>

Question	Part	Author’s response	Marks
10	(a)	7 protons 7 electrons	1 1
	(b)	Electrons carry a negative charge, and protons carry a positive charge. Opposite charges attract.	1 1
	(c)	 **1 mark each for:** 7 protons, 7 neutrons, and 7 electrons Protons and neutrons in the nucleus, electrons around the nucleus Electrons arranged in two shells with 2 and 5 electrons respectively	3
	(d)	2,5	1
11		Atoms absorb energy by promoting an electron to a higher energy level.	1
		The energy levels are fixed, so the amount of energy required to move between them does not vary.	1
12		**Any one of:** The absorption and emission spectra of an element are specific/unique/distinctive. Elements can be identified using their absorption and emission spectra.	1
13	(a)	Atomic absorption	1
	(b)	Atoms absorb energy by promoting an electron to a higher energy shell.	1
		Each dark line corresponds to a particular amount of energy that the atom can absorb; other amounts of energy are not absorbed.	1
		If electrons were not confined to distinct energy levels, atoms could absorb any amount of energy, and all colours would be absorbed equally.	1
	(c)	Reference spectra of magnesium and iron from pure samples are required.	1
		The solar spectrum can be compared to the reference spectra.	1
		If energy is absorbed at all the same wavelengths in both spectra, the element is present.	1

Question	Part	Name	Symbol	Atomic Number	Mass Number	Protons	Electrons	Neutrons	Marks
14		Sodium	$^{23}_{11}\text{Na}$	11	24	**11**	**11**	**13**	1
		Yttrium	$^{89}_{\mathbf{39}}\text{Y}$	**39**	**89**	39	39	**50**	1
		Copper	$^{\mathbf{64}}_{\mathbf{29}}\mathbf{Cu}$	**29**	**64**	**29**	**29**	35	1
		Lead	$^{207}_{\mathbf{82}}\text{Pb}$	**82**	**207**	**82**	**82**	**125**	1
		Thallium	$^{\mathbf{204}}_{\mathbf{81}}\mathbf{Tl}$	81	204	**81**	**81**	**123**	1
		Hydrogen	$^{1}_{1}\text{H}$	**1**	**1**	**1**	**1**	**0**	1
		Carbon	$^{\mathbf{12}}_{\mathbf{6}}\mathbf{C}$	**6**	12	**6**	**6**	**6**	1
		Nitrogen	$^{\mathbf{14}}_{\mathbf{7}}\text{N}$	**7**	**14**	**7**	**7**	7	1
		Barium	$^{\mathbf{137}}_{\mathbf{56}}\mathbf{Ba}$	**56**	137	**56**	**56**	**81**	1
		Calcium	$^{\mathbf{40}}_{\mathbf{20}}\mathbf{Ca}$	**20**	**40**	**20**	**20**	20	1

One mark for each row that is fully correct.

Question	Part	Author's response	Marks
15		Argon-36: 18 protons, 18 electrons, 18 neutrons	1
		Boron-11: 5 protons, 5 electrons, 6 neutrons	1
		Neon-22: 10 protons, 10 electrons, 12 neutrons	1
16	(a)	56 neutrons	1
	(b)	$^{99}_{43}\text{Tc}$	1
17	(a)	Forms of an element that contain differing numbers of neutrons in their nuclei.	1
	(b)	Chemical properties are mostly determined by electron behaviour.	1
		Different isotopes of an element still have the same number of electrons.	1

Question	Part	Isotope	Number of protons	Number of electrons	Number of neutrons	Marks
17	(c)	platinum-190	78	78	112	1
		platinum-192	78	78	114	1
		platinum-194	78	78	116	1
		platinum-195	78	78	117	1
		platinum-196	78	78	118	1
		platinum-198	78	78	120	1

One mark for each row that is fully correct.

Question	Part	Author's response	Marks
18		Electrons in atoms exist at different energy levels. First and foremost, electrons are arranged into main shells, numbered 1, 2, 3, etc.	1
		Each main shell is composed of subshells, ordered *s*, *p*, *d*, *f*, and each electron in each main shell is also assigned to a subshell within that main shell.	1
19		The second period includes elements with their highest energy electrons in the second main shell.	1
		The second main shell does not contain a *d* subshell.	1

Question	Part	Author's response	Marks
20	(a)	$1s^1$	1
	(b)	$1s^22s^2$	1
	(c)	$1s^22s^22p^2$	1
	(d)	$1s^22s^22p^4$	1
	(e)	$1s^22s^22p^5$	1
	(f)	$1s^22s^22p^63s^1$	1
	(g)	$1s^22s^22p^63s^23p^4$	1
	(h)	$1s^22s^22p^63s^23p^6$	1
	(i)	$1s^22s^22p^63s^23p^64s^2$	1
	(j)	$1s^22s^22p^63s^23p^64s^13d^5$	1
	(k)	$1s^22s^22p^63s^23p^64s^13d^{10}$	1
	(l)	$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$	1
	(m)	$1s^22s^22p^63s^23p^64s^23d^{10}4p^4$	1
	(n)	$1s^22s^22p^63s^23p^6$	1
	(o)	$1s^22s^22p^6$	1
	(p)	$1s^22s^22p^63s^23p^6$	1
	(q)	$1s^22s^22p^63s^23p^63d^2$	1
	(r)	$1s^22s^22p^63s^23p^63d^7$	1
	(s)	$1s^22s^22p^63s^23p^63d^{10}$	1
	(t)	$1s^22s^22p^6$	1
21	(a)	$M_{H_2} = (2 \times M_H)$ $= (2 \times 1.01)$ $= 2.02$ g/mol	1
	(b)	$M_{O_3} = (3 \times M_o)$ $= (3 \times 16.00)$ $= 48.00$ g/mol	1
	(c)	$M_{Ne} = 20.18$ *(read from periodic table)*	1
	(d)	$M_{C_8H_{10}N_4O_2} = (8 \times M_C) + (10 \times M_H) + (4 \times M_N) + (2 \times M_O)$ $= (8 \times 12.01) + (10 \times 1.01) + (4 \times 14.00) + (2 \times 16.00)$ $= 194.18$ g/mol	1
	(e)	$M_{LiH} = (1 \times M_{Li}) + (1 \times M_H)$ $= (1 \times 6.94) + (1 \times 1.01)$ $= 7.95$ g/mol	1
	(f)	$M_{NaN_3} = (1 \times M_{Na}) + (3 \times M_N)$ $= (1 \times 22.99) + (3 \times 14.00)$ $= 64.99$ g/mol	1
	(g)	$M_{Al_2(SO4)_3} = (2 \times M_{Al}) + (3 \times M_S) + (12 \times M_O)$ $= (2 \times 26.98) + (3 \times 32.05) + (12 \times 16.00)$ $= 342.11$ g/mol	1
	(h)	$M_{CH_3CH_2CH_2O_H} = M_{C_3H_8O}$ $M_{C_3H_8O} = (3 \times M_C) + (8 \times M_H) + (1 \times M_O)$ $= (3 \times 12.01) + (8 \times 1.01) + (1 \times 16.00)$ $= 60.11$ g/mol	1

Question	Part	Author’s response	Marks
22	(a)	$M_{r(NH_3)} = (1 \times M_N) + (3 \times M_H)$ $= (1 \times 14.00) + (3 \times 1.01)$ $= 17.03$ **(no units)**	1
	(b)	$M_{r(NO_3)} = (1 \times M_N) + (3 \times M_O)$ $= (1 \times 14.00) + (3 \times 16.00)$ $= 62.00$ **(no units)**	1
	(c)	$M_{r(HOCl)} = (1 \times M_H) + (1 \times M_O) + (1 \times M_{Cl})$ $= (1 \times 1.01) + (1 \times 16.00) + (1 \times 35.44)$ $= 52.45$ **(no units)**	1
	(d)	$M_{r(C_4H_{10})} = (4 \times M_C) + (10 \times M_H)$ $= (4 \times 12.01) + (10 \times 1.01)$ $= 58.14$ **(no units)**	1
	(e)	$M_{r(CF_4)} = (1 \times M_C) + (4 \times M_F)$ $= (1 \times 12.01) + (4 \times 19.00)$ $= 88.01$ **(no units)**	1
	(f)	$M_{r(CH_3COCH_3)} = M_{C_3H_6O}$ $M_{C_3H_6O} = (3 \times M_C) + (6 \times M_H) + (1 \times M_O)$ $= (3 \times 12.01) + (6 \times 1.01) + (1 \times 16.00)$ $= 58.09$ **(no units)**	1
23	(a)	$M_{r(CaCl_2)} = (1 \times M_{Ca}) + (2 \times M_{Cl})$ $= (1 \times 40.08) + (2 \times 35.44)$ $= 110.96$ **(no units)**	1
	(b)	$M_{r(Mg_3(PO_4)_2)} = (3 \times M_{Mg}) + (2 \times M_P) + (8 \times M_O)$ $= (3 \times 24.30) + (2 \times 30.97) + (8 \times 16.00)$ $= 262.84$ **(no units)**	1
	(c)	$M_{r(K_2O)} = (2 \times M_K) + (1 \times M_O)$ $= (2 \times 39.10) + (1 \times 16.00)$ $= 94.20$ **(no units)**	1
	(d)	$M_{r(Li_3N)} = (3 \times M_{Li}) + (1 \times M_N)$ $= (3 \times 6.94) + (1 \times 14.00)$ $= 34.82$ **(no units)**	1
	(e)	$M_{r(Sr(OH)_2)} = (1 \times M_{Sr}) + (2 \times M_O) + (8 \times M_H)$ $= (1 \times 87.62) + (2 \times 16.00) + (8 \times 1.01)$ $= 127.70$ **(no units)**	1
	(f)	$M_{r(BaCl_2 \cdot 2H_2O)} = (1 \times M_{Ba}) + (2 \times M_{Cl}) + (4 \times M_H) + (2 \times M_O)$ $= (1 \times 137.3) + (2 \times 35.44) + (4 \times 1.01) + (2 \times 16.00)$ $= 244.22$ **(no units)**	1

Question	Part	Author's response	Marks
24	(a)	$M_{r(C_{12}H_{22}O_{11})} = (12 \times M_C) + (22 \times M_H) + (11 \times M_O)$ $= (12 \times 12.01) + (22 \times 1.01) + (11 \times 16.00)$ $= 342.34$ **(no units)**	1
	(b)	**The mass of one mole is called the molar mass. It is calculated in the same way as the relative molecular mass in Question 24(a), but has units of g/mol.** $M_{r(C_{12}H_{22}O_{11})} = 342.34$ g/mol The question asks for a mass (m), not a molar mass (M). $n = \frac{m}{M} \therefore m = n \times M$ $n \times M = 1$ mol $\times$ 342.34 g/mol = 342.34 g	1 1
	(c)	6.02×10^{23} molecules of sugar (Avogadro's number)	1
	(d)	1 sugar molecule contains 22 hydrogen atoms. Therefore, 1 mol of sugar contains 22 mol of hydrogen atoms.	1
	(e)	1 mol of sugar = 22 mol of hydrogen atoms 22 mol × Avogadro's number = $22 \times 6.02 \times 10^{23}$ $22 \times 6.02 \times 10^{23} =$ 1.32×10^{25} atoms of hydrogen	1
	(f)	$n = \frac{m}{M} \therefore m = n \times M$ $n \times M = 0.5980$ mol × 342.34 g/mol = 204.7 g	1
	(g)	$n = \frac{m}{M}$ $n = \frac{4.2\text{g}}{342.34 \text{ g/mol}}$ = 0.012 mol **or** 12 mmol (millimoles)	1
25		$A_r = (\%\ ^{24}Mg \times A_{^{24}Mg}) + (\%\ ^{25}Mg \times A_{^{25}Mg}) + (\%\ ^{26}Mg \times A_{^{26}Mg})$ $= (0.790 \times 23.985) + (0.100 \times 24.986) + (0.110 \times 25.983)$ ≈ 24.30	1

Question	Atomic Number	Symbol	Name	Classification	Marks
26	2	He	helium	non-metal	1
	30	Zn	**zinc**	**metal**	1
	47	**Ag**	silver	**metal**	1
	32	Ge	**germanium**	**metalloid**	1
	15	**P**	phosphorus	**non-metal**	1
	5	**B**	**boron**	**metalloid**	1
	33	**As**	arsenic	**metalloid**	1
	19	**K**	**potassium**	**metal**	1
	79	Au	**gold**	**Metal**	1

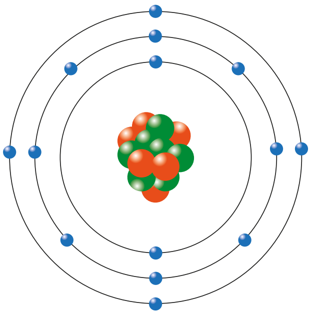
One mark for each row that is fully correct.

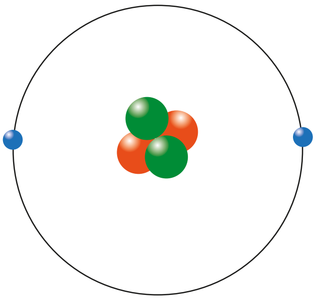
Question	Part	Author's response					Marks
		Atomic Number	**Symbol**	**Name**	**Classification**	**Block**	
27		54	Xe	xenon	noble gas	p-block	1
		41	**Nb**	**niobium**	**transition metal**	**d-block**	1
		12	**Mg**	**magnesium**	**alkaline earth metal**	**s-block**	1
		82	**Pb**	lead	**post-transition metal**	**p-block**	1
		9	F	**fluorine**	**halogen**	**p-block**	1
		59	**Pr**	praseodymium	**lanthanide**	**f-block**	1
		92	U	**uranium**	**actinide**	**f-block**	1
		37	**Rb**	**rubidium**	**alkali metal**	**s-block**	1
		36	**Kr**	krypton	**noble gas**	**p-block**	1
		One mark for each row that is fully correct.					

Question	Part	Electron Configuration	Period Number	Group Number	Element Name	Marks
28		$1s^22s^22p^2$	**2**	**XIV (main group IV)**	**carbon**	1
		$1s^22s^22p^63s^23p^64s^23d^6$	**4**	**VIII**	**iron**	1
		$1s^22s^22p^63s^23p^6$	**3**	**XVIII (main group VIII)**	**argon**	1
		$1s^22s^22p^63s^23p^64s^13d^5$	**4**	**VI**	**chromium**	1
		$1s^22s^2$	**2**	**II**	**beryllium**	1
		$1s^22s^22p^4$	**2**	**XVI (main group VI)**	**oxygen**	1
		$1s^22s^22p^63s^1$	**3**	**I**	**sodium**	1
		$1s^22s^22p^63s^23p^5$	**3**	**XVII (main group VII)**	**chlorine**	1
		$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$	**4**	**XIII (main group XVIII)**	**Gallium**	1
		One mark for each row that is fully correct.				

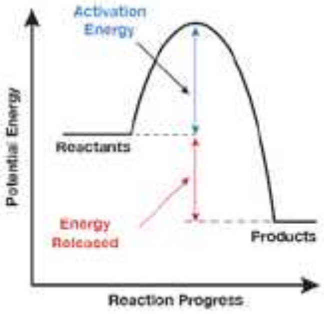
Question	Part	Author's response	Marks
29	(a)	Fluorine has a greater ability to attract electrons to itself than oxygen.	1
	(b)	High electronegativity	1
	(c)	Fluorine	1
	(d)	Oxygen	1
	(e)	Atomic radius decreases from left to right across a period. Oxygen has fewer protons to attract its outermost electrons to the nucleus.	1 1

Summary Test 1: Materials and their Atoms

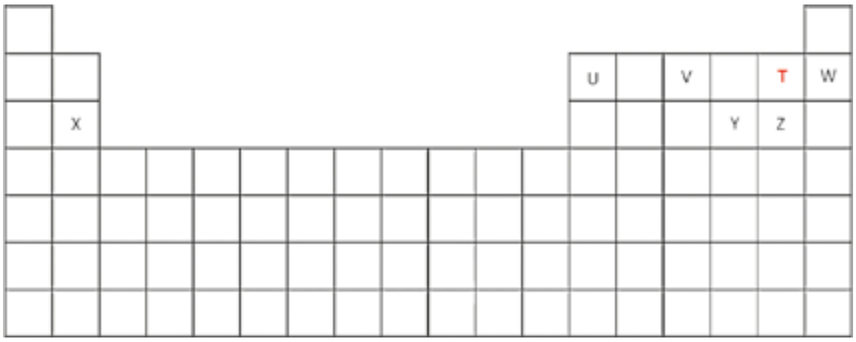
Question	Part	Author's response	Marks
1	(a)	Filtration	1
	(b)	The mixture is passed through a filter.	1
		The sand (solid) is trapped, but the water (liquid) passes through.	1
	(c)	 **1 mark each for:** 14 protons, 14 neutrons, and 14 electrons Protons and neutrons in the nucleus, electrons around the nucleus Electrons arranged in three shells with configuration 2,8,4	3
	(d)	Metalloid	1
	(e)	28.08 **(no units)**	1
	(f)	$M_{r(SiO2)} = (1 \times M_{Si}) + (2 \times M_O)$ $= (1 \times 28.08) + (2 \times 16.00)$ $= 60.08$ **(no units)**	
	(g)	Silicon makes up $\frac{28.08}{60.08} = 47\%$ of silicon dioxide. The remaining 53% $\left(\frac{32.00}{60.08}\right)$ is made up of oxygen. Oxygen contributes more mass overall to silicon dioxide.	1
2	(a)	Evaporation	1
	(b)	The mixture is heated in an evaporating dish.	1
		The solvent has a lower boiling point and evaporates, leaving the solid solute in the dish.	1
	(c)	Filtration cannot separate dissolved substances from solvents.	1
		Filters have pores small enough to block solid particles, but large enough to allow individual molecules. Dissolved substances have broken apart into individual molecules and are able to pass through a filter.	1
	(d)	$1s^2 2s^2 2p^6$	1
	(e)	$1s^2 2s^2 2p^6$	1
3	(a)	d-block	1
	(b)	High electrical conductivity	1
	(c)	Electricity would be conducted into the handle.	1
		Anyone who touched the handle could be electrocuted.	1
	(d)	High melting point	1
		Electrical conductivity with high resistance	1
	(e)	The wire in the cord would heat up as the kettle was operating.	1
		Plus, any one of: The wire could break. The plastic coating on the wire could melt. Anyone who touched the cord could be burnt.	1

Question	Part	Author's response	Marks
4	(a)	Quantum dots are smaller than 100 nm.	1
	(b)	QDLED displays can produce better quality images. They may also allow for new applications, such as curved displays.	1 1
	(c)	Isotopes are forms of an element that contain differing numbers of neutrons in their nuclei. For example, selenium-74 has 40 neutrons, while selenium-76 has 42 neutrons.	1 1
	(d)	The relative atomic mass is a weighted average which takes into account the relative amounts of each isotope on Earth.	1
5	(a)	 **1 mark each for:** 2 protons, 2 neutrons, and 2 electrons Protons and neutrons in the nucleus, electrons around the nucleus Electrons in one shell	3
	(b)	Non-metal	1
	(c)	Atoms can absorb energy by promoting an electron to a higher energy shell. The dark line corresponds to a particular amount of energy that the atom absorbed; other amounts of energy are not absorbed.	1 1
	(d)	The absorption spectra of each element are unique. The dark line was not present in the spectra of known elements, so must have been caused by the presence of a previously undiscovered element.	1 1
6	(a)	2,8	1
	(b)	$1s^2 2s^2 2p^6$	1
	(c)	The colour produced by neon is composed of the many wavelengths that neon can emit. These wavelengths do not vary, so the overall colour also does not vary.	1 1
	(d)	In ground-state electron configurations, electrons exist at the lowest available energy level. In the excited-state electron configuration given, an electron has been promoted from the 2p subshell to the 4s subshell.	1 1
7	(a)	$M_{r(NaCl)} = (1 \times M_{Na}) + (1 \times M_{Cl})$ $= (1 \times 22.99) + (1 \times 35.44)$ = 58.43 **(no units)**	1
	(b)	**The mass of one mole is called the molar mass. It is calculated in the same way as the relative molecular mass in Question 7(a), but has units of g/mol.** $M_{NaCl} = 58.43$ g/mol	1
	(c)	$n = \frac{m}{M}$ $n = \frac{3.5g}{58.43\ g/mol}$ = 0.06 mol **or** 60 mmol (millimoles)	1
	(d)	1 formula unit of sodium chloride contains 1 sodium atom. Therefore, 1 mol of sodium chloride contains 1 mol of sodium atoms. Therefore, 0.06 mol of sodium chloride contains 0.06 mol of sodium atoms. 0.06 mol = $0.06 \times 6.02 \times 10^{23}$ particles = 3.60×10^{22} atoms of sodium	1

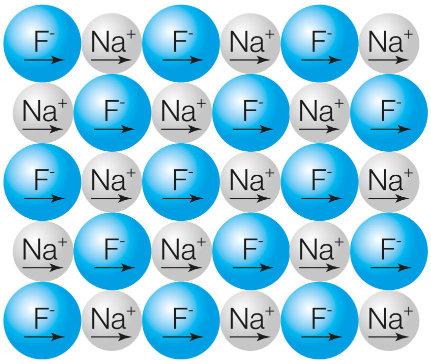
Topic 2: Combining Atoms

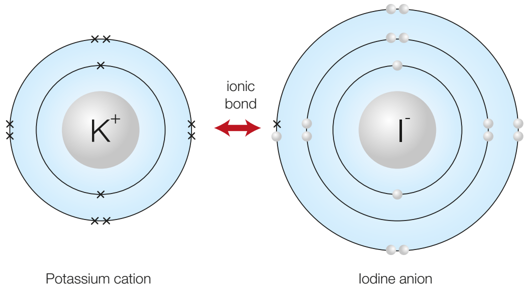
Question	Part	Author's response	Marks
1	(a)	1+	1
	(b)	When valence electrons of atoms interact, metal atoms such as potassium will release their valence electrons. For potassium this is one valence electron. This results in a positively charged ion referred to as a cation.	1 1
	(c)	K^+	1
2	(a)	6	1
	(b)	Oxygen	1
	(c)	VI	1
	(d)	When valence electrons of atoms interact, non-metal atoms such as atom X will gain valence electrons to complete its valence shell. Atom X will gain two electrons to form a negatively charged ion referred to as an anion.	1 1
	(e)	$1s^22s^22p^6$	1
3	(a)	Exothermic reactions occur when a greater quantity of energy is released when forming new bonds than what is absorbed to break the bonds of the reactants. Fluorine gaining an electron is an exothermic process as it results in a release 328 kJ/mol of energy as the atom completes its octet electron configuration.	1 1
	(b)	 The reaction between fluorine and hydrogen releases 271 kJ/mol of energy. As the energy of the products being lower than the energy of the reactants the reaction is exothermic.	4 1 1
	(c)	Activation energy is the minimum amount of energy that must be provided to the reactant molecules of fluorine and hydrogen to result in a chemical reaction. The activation energy (E_A) of a reaction is measured in J/mol.	1 1
4	(a)	Titanium atoms release their two valence electrons resulting in a "sea" of delocalised electrons forming positively charged cations. The attraction between the positive nuclei and "sea" of delocalised electrons results in the formation of the metallic lattice structure.	1 1
	(b)	The valence electrons of each metallic atom are shared by more than one neighbouring atom. The valence shells of these adjacent atoms overlap, allowing electrons to move freely within these valence shells, "delocalised" from their original atom.	1 1
	(c)	The atomic radius of titanium would be larger than that of zinc, decreasing the strength of attraction between the valence electrons and positive nucleus. Metallic bonds are weaker between atoms that have a larger atomic radius.	1 1

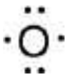
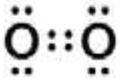
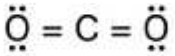
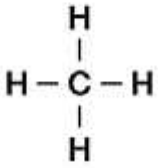
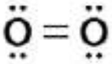
Question	Part	Author's response	Marks
5	(a)	Strong electrostatic forces of attraction hold the positively charged nuclei and the "sea" of delocalised electrons together. This results in a formed material that is comprised of a tightly packed, dense and 3D regular lattice.	1 1
	(b)	Strong electrostatic forces of attraction hold barium's positively charged nuclei and "sea" of delocalised electrons together. Sufficient energy must be supplied to overcome barium's metallic bonds resulting in a high melting point.	1 1
	(c)	Malleability	1
	(d)	Metals such as barium are electrical conductors as they have charged particles that are free to move. Barium's delocalised electrons will move towards a positive electrode and away from a negative electrode in an electric circuit.	1 1
	(e)	Any two of: Electrical conductor Hard High melting point Malleable	1
	(f)	Electrostatic	1
	(g)	Aluminium has a greater positive charge and smaller atomic radius than barium. Metallic bonds are stronger between atoms which have a greater charge and smaller radius.	1 1
		Therefore aluminium will have the greater melting point.	1
6	(a)	Electrical and/or thermal conductivity. Lustrous	1 1
	(b)	Electrical and thermal conductivity result from the presence of delocalised electrons. Lustre also results from the interaction of photons and delocalised electrons. To share these properties, graphite's structure must also include the presence of delocalised electrons.	1 1 1

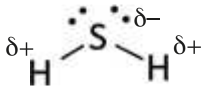
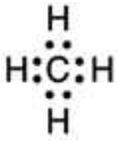
<table>
<tr><th>Question</th><th>Part</th><th>Author's response</th><th>Marks</th></tr>
<tr><td rowspan="7">7</td><td>(a)</td><td>
<table>
<tr><th>Atom</th><th>Formula and charge</th></tr>
<tr><td>V</td><td>V^{3-}</td></tr>
<tr><td>X</td><td>X^{2+}</td></tr>
<tr><td>Z</td><td>Z^{-}</td></tr>
</table>
</td><td>3</td></tr>
<tr><td>(b)</td><td>
<table>
<tr><th>Atom</th><th>Anion / Cation</th></tr>
<tr><td>U</td><td>Cation</td></tr>
<tr><td>X</td><td>Cation</td></tr>
<tr><td>Y</td><td>Anion</td></tr>
</table>
</td><td>3</td></tr>
<tr><td>(c)</td><td>

*any Group VII element is acceptable</td><td>1</td></tr>
<tr><td>(d)</td><td>As the released electrons are negatively charged, the metal ion gains an overall positive charge equal in magnitude to the number of electrons lost.
This results in a positively charged particle as the number of positively charged protons now exceeds the number of negatively charged electrons.</td><td>1
1</td></tr>
<tr><td>(e) (i)</td><td>
<table>
<tr><th>Atom</th><th>Electron Configuration</th></tr>
<tr><td>V</td><td>$1s^2 2s^2 2p^3$</td></tr>
<tr><td>X</td><td>$1s^2 2s^2 2p^6 3s^2$</td></tr>
</table>
</td><td>2</td></tr>
<tr><td>(e) (ii)</td><td>Atom V is in group VIII and will gain three electrons to complete it's valence shell.
Atom X is in group II and will release two electrons to empty it's valence shell.
The formed ion's electron configuration mirror the completed octet electron configuration of atom W.</td><td>1
1
1</td></tr>
<tr><td rowspan="2">8</td><td>(a)</td><td>
<table>
<tr><th>Everyday Substance</th><th>Cation</th><th>Anion</th><th>Ionic Compound</th></tr>
<tr><td>Table Salt</td><td>Sodium</td><td>Chlorine</td><td>sodium chloride</td></tr>
<tr><td>Bath Salt</td><td>Magnesium</td><td>Chlorine</td><td>magnesium chloride</td></tr>
<tr><td>Rust</td><td>Iron</td><td>Oxygen</td><td>iron oxide</td></tr>
<tr><td>Toothpaste ingredient</td><td>Sodium</td><td>Fluorine</td><td>sodium fluoride</td></tr>
<tr><td>Rock Salt</td><td>Calcium</td><td>Chlorine</td><td>calcium chloride</td></tr>
</table>
</td><td>1
1
1
1
1</td></tr>
<tr><td>(b)</td><td>Calcium atoms release two valence electrons to chlorine.
The valence electrons are gained by two atoms of chlorine.
An ionic bond forms between the calcium cation and chloride anions due to the electrostatic attraction between the oppositely charged ions forming calcium chloride.</td><td>1
1
1</td></tr>
</table>

Question	Part	Author's response		Marks
8	(c)	Calcium is in group II and subsequently releases two valence electrons. Sodium is in group I releasing one valence electron in chemical reactions. Chlorine is in group VII thus requires only valence electron to complete its valence shell, therefore two atoms of chlorine are required when forming calcium chloride.		1 1 1
	(d)			3
9	(a)	**Ionic Compound**	**Empirical Formula**	
		sodium phosphide	Na_3P	1
		aluminium sulfide	Al_2S_3	1
		beryllium chloride	$BeCl_2$	1
		gallium nitride	GaN	1
		calcium oxide	CaO	1
		lithium bromide	LiBr	1
		magnesium fluoride	MgF_2	1
		potassium phosphide	K_3P	1
	(b)	**Ionic Compound**	**Empirical Formula**	
		lead(II) sulfide	PbS	1
		copper(I) arsenide	Cu_3As	1
		iron(III) oxide	Fe_2O_3	1
		iron(II) bromide	$FeBr_2$	1
		mercury(II) oxide	HbO	1
		chromium(III) chloride	$CrCl_3$	1
		tin(IV) chloride	$SnCl_4$	1
		iron(II) sulfide	FeS	1

<table>
<tr><th>Question</th><th>Part</th><th>Author's response</th><th>Marks</th></tr>
<tr><td>9</td><td>(c)</td><td><table>
<tr><th>Ionic Compound</th><th>Empirical Formula</th></tr>
<tr><td>magnesium hydroxide</td><td>$Mg(OH)_2$</td></tr>
<tr><td>sodium nitrate</td><td>$NaNO_3$</td></tr>
<tr><td>vanadium(V) phosphate</td><td>$V_3(PO_4)_5$</td></tr>
<tr><td>potassium sulfate</td><td>K_2SO_4</td></tr>
<tr><td>calcium nitrite</td><td>$Ca(NO_2)_2$</td></tr>
<tr><td>sodium bicarbonate</td><td>Na_2CO_3</td></tr>
<tr><td>copper(I) carbonate</td><td>Cu_2CO_3</td></tr>
<tr><td>lead(II) sulfite</td><td>$PbSO_3$</td></tr>
<tr><td>beryllium acetate</td><td>$BeCH_3COO$</td></tr>
<tr><td>ammonium hydroxide</td><td>NH_4OH</td></tr>
</table></td><td>1
1
1
1
1
1
1
1
1
1</td></tr>
<tr><td>10</td><td>(a)</td><td>Potassium atoms release one valence electron to iodine.
The valence electron is gained by an atoms of iodine.
An ionic bond forms between the potassium cation and iodide anion due to the electrostatic attraction between the oppositely charged ions forming potassium iodide. The repeating 3D lattice forms a solid compound.</td><td>1
1
1</td></tr>
<tr><td></td><td>(b)</td><td>
</td><td>3</td></tr>
<tr><td></td><td>(c)</td><td>KI</td><td>1</td></tr>
<tr><td></td><td>(d)</td><td>Strong electrostatic forces of attraction hold the positively charged potassium cations and negatively charged iodide ions together.
Sufficient energy must be supplied to overcome potassium iodide's ionic bonds resulting in a high melting point.</td><td>1
1</td></tr>
<tr><td></td><td>(e)</td><td>Soluble ionic compounds such as potassium iodide dissolve to release 1+ ions that can carry charge.
Soluble ionic compounds that release ions of greater charge magnitude (such as $MgSO_4$) will carry more amperes of current. Potassium iodide will have comparatively lower conductivity.</td><td>1
1</td></tr>
<tr><td></td><td>(f)</td><td>Ionic compounds such as potassium iodide are brittle substances.
When exposed to mechanical stress they will fracture and are termed substances of low compressive strength.</td><td>1
1</td></tr>
<tr><td>11</td><td>(a)</td><td>When sufficient energy is applied, potassium iodide will melt.
This breaks the ionic bonds, releasing the ions from the lattice.</td><td>1
1</td></tr>
<tr><td></td><td>(b)</td><td>Free ions are required for a substance to conduct electricity.
With the ionic lattice broken in molten sodium chloride, the ions are free to carry change.
In solid sodium chloride the ions are fixed within the lattice, therefore cannot carry charge.</td><td>1
1
1</td></tr>
</table>

<table>
<tr><th>Question</th><th>Part</th><th>Author’s response</th><th>Marks</th></tr>
<tr><td>12</td><td>(a)</td><td></td><td>2</td></tr>
<tr><td></td><td>(b)</td><td>Both oxygen atoms possess two non-bonding pairs of electrons.
Two atoms of oxygen can share four electrons, with the electrostatic attraction between the two bonding electron pairs and positive nuclei of the atoms resulting in the formation of a double bond.</td><td>1
1</td></tr>
<tr><td></td><td>(c)</td><td></td><td>2</td></tr>
<tr><td></td><td>(d) (i)</td><td>2NO + O_2 → $2NO_2$</td><td>1</td></tr>
<tr><td></td><td>(d) (ii)</td><td>The nitrogen dioxide contains both a single and double covalent bond.
A single covalent bond involves two bonding electrons, a double covalent bond involves four bonding electrons.</td><td>1
1</td></tr>
<tr><td></td><td>(d) (iii)</td><td>Electronegativity is the tendency of an atom to attract bonding electrons towards itself.</td><td>1</td></tr>
<tr><td></td><td>(d) (iv)</td><td>Electronegativity increases across a period, and up the periodic groups. Oxygen is in group VI, period II compared to nitrogen in group V, period II resulting in nitrogen’s lower electronegativity.
This is because oxygen has a greater positive charge in the nucleus and a smaller atomic radius, so attracts the bonding electrons more strongly.</td><td>1
1</td></tr>
<tr><td></td><td></td><td>
<table>
<tr><th>Name</th><th>Chemical formula</th><th>Lewis structure diagram</th></tr>
<tr><td>Carbon dioxide</td><td>CO_2</td><td></td></tr>
<tr><td>Methane</td><td>CH_4</td><td></td></tr>
<tr><td>Hydrogen</td><td>H_2</td><td></td></tr>
<tr><td>Oxygen</td><td>O_2</td><td></td></tr>
</table>
</td><td>2
2
2
2</td></tr>
</table>

Question	Part	Author's response	Marks
13	(a)		2
	(b)		1
	(c)	Polar covalent bonds arise between atoms of differing electronegativities. The difference in the atom's tendency to attract bonding electrons results in an unequal sharing of the electron pair, inducing a partial dipole.	1 1
	(d) (i)		2
	(d) (ii)	Non-polar covalent bonds. The bonding electrons have equal attraction to the positive nuclei due to their nil difference in electronegativity resulting in no partial dipole.	1
	(e) (i)	H — H	2
	(e) (ii)	Homonuclear diatomic molecules contain two atoms of the same element. The atoms share the same electronegativity as the bonding electrons have equal attraction to the positive nuclei resulting in no partial dipole.	1 1

Question	Systematic molecular name	Molecular formula	Marks
14	chlorine dioxide	ClO_2	1
	oxygen difluoride	OF_2	1
	sulfur tetrachloride	SCl_4	1
	diphosphorus pentoxide	P_2O_5	1
	sulfur dichloride	SCl_2	1
	bromine pentafluoride	BrF_5	1
	dinitrogen trioxide	N_2O_3	1
	xenon trioxide	XeO_3	1
	sulfur hexafluoride	SF_6	1
	iodine dichloride	ICl_2	1
	carbon monoxide	CO	1
	dinitrogen pentasulfide	N_2S_5	1

Question	Part	Author's response		Marks
		Molecular formula	**Systematic molecular name**	
15		ClO_2	chlorine dioxide	1
		P_4S_5	tetraphosphorus pentasulfide	1
		N_2S	dinitrogen sulphide	1
		SCl_4	sulfur tetrachloride	1
		B_2Si	diboron silicide	1
		NF_3	nitrogen trifluoride	1
		HI	hydrogen iodide	1
		PI_3	phosphorus triiodide	1
		IF_5	iodine pentafluoride	1
		N_2O_3	dinitrogen trioxide	1
		SO_3	sulfur trioxide	1
		CS_2	carbon disulfide	1
16	(a)	Electrical conductivity.		1
		In graphite's network covalent structure only three out of the four valence electrons in the carbon atoms are involved in covalent bonding.		1
		The fourth valence electron is delocalised, and can thus move freely between the graphite layers, acting as a current carrier.		1
	(b)	Covalent networks are similar in structure to ionic compounds forming a repeating three-dimensional structure.		1
		In diamond, each carbon shares electrons with four adjacent carbon atoms to form a covalent network in a tetrahedral arrangement.		1
		In graphite each carbon atom covalently bonds to three adjacent carbon atoms to form layers with a hexagonal arrangement.		1
		The tetrahedral arrangement results in a very hard substance, whereas the hexagonal arrangement results in weaker interactions producing a substance that can be broken and scratched.		1
	(c)	Graphite is a covalent network solid consisting of a repeating three-dimensional structure.		1
		Methane is a covalent molecule.		1
		Methane molecules are discrete units held together by weaker secondary interactions, compared to the primary covalent bonds which bond the covalent lattice structure of graphite.		1
		The strong repeating primary covalent bonds in graphite produce a substance of solid of higher melting point than the gaseous methane.		1
17	(a)	Each silicon atom forms four covalent bonds with the adjacent oxygen atoms. There are two oxygen atoms to every silicon atom.		1
		Its bonding arrangement results in a similar structure to diamond with no free electrons and a repeating three-dimensional structure resulting in a solid of extremely high melting point.		1
	(b)	Non-conductive.		1
		As silica's bonding arrangement does not have any free electrons, the substance is unable to carry electricity, allowing it to be used in canning of food and drinks.		1

Question	Part	Author's response	Marks
18	(a)	Methionine – $C_5H_{11}NO_2S$ M = 75.07 g mol^{-1} $n = \frac{m}{M}$ $n = \frac{2}{75.07}$ $n = 2.67 \times 10^{-2}$ mol $N = n \times N_A$ $N = 2.67 \times 10^{-2} \times 6.02 \times 10^{23}$ $N = 1.60 \times 10^{22}$ molecules present in the sample	1 1 1 1
	(b)	Glutamic acid – $C_5H_9NO_4$ M = 147.13 g mol^{-1} $n = \frac{m}{M}$ $n = \frac{2}{147.13}$ $n = 1.36 \times 10^{-2}$ mol $N = n \times N_A$ $N = 1.36 \times 10^{-2} \times 6.02 \times 10^{23}$ $N = 8.18 \times 10^{21}$ molecules present in the sample	1 1 1 1
	(c)	Tryptophan – $C_{11}H_{12}N_2O_2$ M = 204.23 g mol^{-1} $n = \frac{m}{M}$ $n = \frac{2}{204.23}$ $n = 9.80 \times 10^{-3}$ mol $N = n \times N_A$ $N = 9.80 \times 10^{-3} \times 6.02 \times 10^{23}$ $N = 5.90 \times 10^{21}$ atoms present in the sample	1 1 1 1
	(d)	Valine – $C_5H_{11}NO_2$ M = 117.15 g mol^{-1} $n = \frac{m}{M}$ $n = \frac{2}{117.15}$ $n = 1.70 \times 10^{-2}$ mol $N = n \times N_A$ $N = 1.70 \times 10^{-2} \times 6.02 \times 10^{23}$ $N = 1.03 \times 10^{22}$ molecules present in the sample	1 1 1 1

Question	Part	Author's response	Marks
18	(e)	Theanine – $C_7H_{14}N_2O_3$ $M = 174.2\ g\ mol^{-1}$ $n = \frac{m}{M}$ $n = \frac{2}{174.2}$ $n = 1.15 \times 10^{-2}$ mol $N = n \times N_A$ $N = 1.15 \times 10^{-2} \times 6.02 \times 10^{23}$ $N = 6.91 \times 10^{21}$ molecules present in the sample	1 1 1 1
	(f)	Glycine – $C_2H_5NO_2$ $M = 75.07\ g\ mol^{-1}$ $n = \frac{m}{M}$ $n = \frac{2}{75.07}$ $n = 2.66 \times 10^{-2}$ mol $N = n \times N_A$ $N = 2.66 \times 10^{-2} \times 6.02 \times 10^{23}$ $N = 1.60 \times 10^{22}$ molecules present in the sample	1 1 1 1
19	(a)	$M = 63.55\ g\ mol^{-1}$ $n = \frac{m}{M}$ $n = \frac{7.88}{63.55}$ $n = 1.24 \times 10^{-1}$ mol	1 1
	(b)	$N = n \times N_A$ $N = 1.24 \times 10^{-1} \times 6.02 \times 10^{23}$ $N = 7.46 \times 10^{22}$ molecules present in the sample	1 1
20	(a)	$M = (4 \times M_C) + (6 \times M_H) + (6 \times M_O)$ $M = (4 \times 12.01) + (6 \times 1.008) + (6 \times 16)$ $M = 150.088\ g\ mol^{-1}$	1
	(b)	$n = \frac{m}{M}$ $n = \frac{125}{150.088}$ $n = 8.33 \times 10^{-1}$ mol	1 1
	(c)	$N = n \times N_A$ $N = 8.33 \times 10^{-1} \times 6.02 \times 10^{23}$ $N = 5.01 \times 10^{23}$ molecules present in the sample	1 1

Question	Part	Author's response	Marks
21	(a)	$M = M_{Fe} + M_{O}$ $M = 71.85\ \text{g mol}^{-1}$ percentage composition $\% = \frac{M_{element} \times N_{atoms\ of\ that\ element}}{M_{compound}} \times 100$ $\% = \frac{55.85 \times 1}{71.85} \times 100$ $\% = 77.73\%$	1 1
	(b)	$M = (2 \times M_{Fe}) + (3 \times M_{O})$ $M = 159.70\ \text{g mol}^{-1}$ percentage composition $\% = \frac{M_{element} \times N_{atoms\ of\ that\ element}}{M_{compound}} \times 100$ $\% = \frac{55.85 \times 2}{159.70} \times 100$ $\% = 69.94\%$	1 1
	(c)	FeO	1
22		$M = (9 \times M_{C}) + (8 \times M_{H}) + (4 \times M_{O})$ $M = (9 \times 12.01) + (8 \times 1.008) + (4 \times 16.00)$ $M = 180.154\ \text{g mol}^{-1}$ **Carbon percent composition** percentage composition $\% = \frac{M_{element} \times N_{atoms\ of\ that\ element}}{M_{compound}} \times 100$ $\% = \frac{12.01 \times 9}{180.154} \times 100$ $\% = 60.00\%$ **Hydrogen percent composition** percentage composition $\% = \frac{1.008 \times 8}{180.154} \times 100$ $\% = 4.48\%$ **Oxygen percent composition** percentage composition $\% = \frac{M_{element} \times N_{atoms\ of\ that\ element}}{M_{compound}} \times 100$ $\% = \frac{16.00 \times 4}{180.154} \times 100$ $\% = 35.52\%$	1 1 1 1 1 1 1

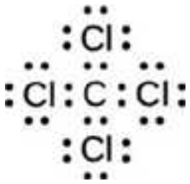
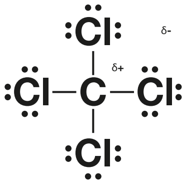
Question	Part	Author’s response	Marks
23		m_c = % composition × 100 $m_c = 0.2729 \times 100$ $m_c = 27.29$ g m_O = % composition × 100 $m_O = 0.7271 \times 100$ $m_O = 72.71$ g $n_C = \frac{m}{M}$ $n_C = \frac{27.29}{12.01}$ $n_C = 2.27$ mol $n_O = \frac{m}{M}$ $n_O = \frac{72.71}{16}$ $n_O = 4.54$ mol ∴ $C_{2.27}O_{4.54}$ $C = \frac{2.27}{2.27}$ $O = \frac{4.54}{2.27}$ Therefore, the empirical formula of the compound is CO_2	1 1 1 1 1 1 1 1
24		m_c = % composition × 100 $m_c = 0.7402 \times 100$ $m_c = 74.02$ g m_H = % composition × 100 $m_H = 0.0871 \times 100$ $m_H = 8.710$ g m_N = % composition × 100 $m_N = 0.1727 \times 100$ $m_N = 17.27$ g $n_C = \frac{m}{M}$ $n_C = \frac{74.02}{12.01}$ $n_C = 6.16$ mol $n_H = \frac{m}{M}$ $n_H = \frac{8.710}{1.008}$ $n_H = 8.64$ mol $n_N = \frac{m}{M}$ $n_N = \frac{17.27}{14.01}$ $n_N = 1.23$ mol ∴ $C_{6.16}H_{8.64}N_{1.23}$ $C = \frac{6.16}{1.23}$ $H = \frac{8.64}{1.23}$ $N = \frac{1.23}{1.23}$ Therefore, the empirical formula of the compound is C_5H_7N	

Question	Part	Author's response	Marks
25		$n_{Fe} = \frac{m}{M}$ $n_{Fe} = \frac{69.95}{55.85}$ $n_{Fe} = 1.25$ mol $n_O = \frac{m}{M}$ $n_O = \frac{30.06}{1.89}$ $n_O = 1.89$ mol $\therefore Fe_{1.25}O_{1.89}$ $Fe = \frac{1.25}{1.25} \quad O = \frac{1.89}{1.25}$ $\therefore FeO_{1.5} \times 2$ $\therefore Fe_2O_3$	

Summary Test 2: Combining Atoms

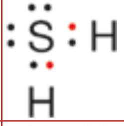
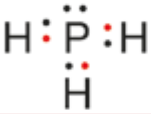
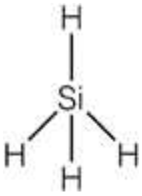
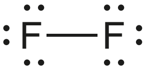
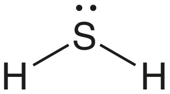
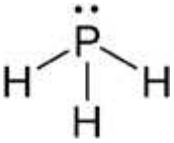
<table>
<tr><th>Question</th><th>Part</th><th>Author's response</th><th>Marks</th></tr>
<tr><td rowspan="6">1</td><td>(a)</td><td>Aluminium atoms release their three valence electrons resulting in a "sea" of delocalised electrons forming positively charged cations.
The attraction between the positive nuclei and "sea" of delocalised electrons results in the formation of the metallic lattice structure.</td><td>1
1</td></tr>
<tr><td>(b)</td><td><table>
<tr><th>Metal</th><th>Group</th><th>Period</th><th>Block</th></tr>
<tr><td>lithium</td><td>I</td><td>2</td><td>s</td></tr>
<tr><td>potassium</td><td>I</td><td>4</td><td>s</td></tr>
<tr><td>aluminium</td><td>III</td><td>3</td><td>p</td></tr>
</table></td><td>3
3
3</td></tr>
<tr><td>(c)</td><td><table>
<tr><th>Metal</th><th>Electron Configuration</th></tr>
<tr><td>lithium</td><td>$1s^2 2s^1$</td></tr>
<tr><td>potassium</td><td>$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$</td></tr>
<tr><td>aluminium</td><td>$1s^2 2s^2 2p^6 3s^2 3p^1$</td></tr>
</table></td><td>2
2
2</td></tr>
<tr><td>(d)</td><td>Both lithium and potassium are group I metals with a valency of one.
Aluminium is a group III metal and thus has a valence of three.</td><td></td></tr>
<tr><td>(e) (i)</td><td>Lithium has a smaller atomic radius compared to potassium.
This increases the attractive force between the "sea" of delocalised electrons and the positive nuclei resulting in a higher boiling point.</td><td></td></tr>
<tr><td>(e) (ii)</td><td>Aluminium is a group III metal, possessing a +3 cation when bonding. This is a greater charge magnitude than that of lithium and potassium, group I metals which produces a +2 cation.
The aluminium cation has greater electrostatic attraction for delocalised electrons increasing the amount of energy required to overcome the attraction to melt and subsequently boil it.</td><td></td></tr>
</table>

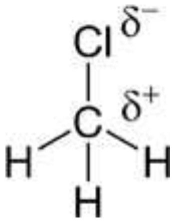
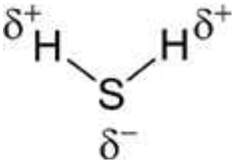
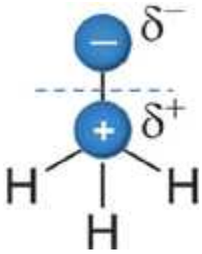
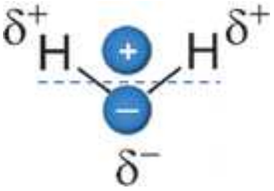
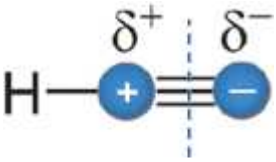
Question	Part	Author's response	Marks
1	(f) (i)	Hard, thermal conductor and high melting point.	1 + 1 + 1
	(f) (ii)	In an alloy the atoms in the lattice structure are of different sizes. The difference in atomic size distorts the regular lattice structure. This results in alloys requiring a greater quantity of force to allow the layers to slide over one another, therefore alloys are often harder than the pure metals they are comprised of.	1 1
	(f) (iii)	Metals such as aluminium and stainless steel are electrical conductors as they have charged particles that are free to move. Delocalised electrons will move towards a positive electrode and away from a negative electrode in an electric circuit.	1 1
2	(a) (i)	2	
	(a) (ii)	When valence electrons of atoms interact, metal atoms such as magnesium will release two valence electrons. This results in a positively charged ion referred to as a cation.	1 1
	(a) (iii)	Magnesium releases two valence electrons from its *3s* subshell. This will reduce the atomic radium of the formed magnesium ion.	1 1
	(b)	Magnesium atoms release two valence electrons to chlorine. The valence electrons are gained by two atoms of chlorine. An ionic bond forms between the magnesium cation and chloride anion due to the electrostatic attraction between the oppositely charged ions forming magnesium chloride. The repeating 3D lattice forms a solid compound.	1 1 1
	(c) (i)	$MgSO_4$	1
	(c) (ii)	Polyatomic ions such as "sulfate" are those composed of more than one atom. Polyatomic ions arise when a covalent molecule undergoes an electron exchange.	1 1
	(d) (i)	When ionic compounds such as bath salts are exposed to compressive or mechanical stress, the layers of ions shift. This forces the ions of like charges close to one another. The induced repulsion causes the solid to shatter across a smooth plane of the material.	1 1
	(d) (ii)	The strength of ionic attraction is increased when the ions contained in the lattice are of greater charge magnitude. This results in a greater requirement of applied energy to separate them, increasing the compounds melting point. Thus magnesium chloride ($MgCl_2$) will have a higher melting point that magnesium sulfate ($MgSO_4$).	1
	(d) (iii)	The strength of ionic attraction is increased when the ions contained in the lattice are of greater charge magnitude. When exposed to mechanical stress they will fracture and are termed substances of low compressive strength. Thus magnesium chloride ($MgCl_2$) will have a higher melting point that magnesium sulfate ($MgSO_4$).	1 1 1
	(e) (i)	Electrolytes are compounds which produce ions when dissolved into an aqueous solution with water (if soluble) or when molten	1 + 1
	(e) ii)	Free ions are required for a substance to conduct electricity. With the ionic lattice broken in molten sodium chloride, the ions are free to carry change. In solid sodium chloride the ions are fixed within the lattice, therefore cannot carry charge.	1 1 1

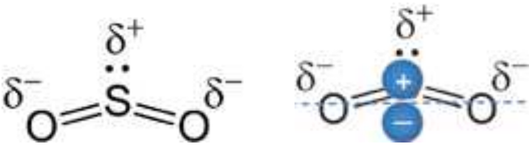
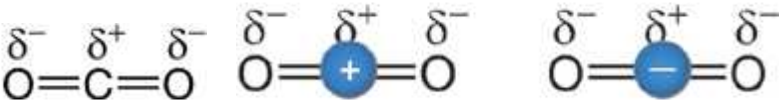
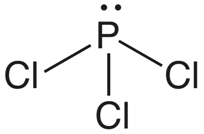
Question	Part	Author's response	Marks
3	(a)	Carbon: $1s^2 2s^2 2p^2$	2
		Chlorine: $1s^2 2s^2 2p^6 3s^2 3p^5$	2
	(b)	CCl_4	1
	(c)		2
	(d) (i)	Electronegativity increases across a period, and up the periodic groups. Chlorine is in **group VII**, period III compared to carbon in **group IV**, period II resulting in carbon's lower electronegativity.	
	(d) (ii)	A single covalent bond involves two atoms sharing one pair of electrons to complete their octet. Group VII elements such as chlorine have incomplete valence shells with 7 electrons. By combining with an atom, both atoms complete their valence shell and attain octet stability. The bonding electrons are electrostatically attracted to the positive nuclei of both atoms, covalently bonding the atoms together.	
	(d) (iii)	Molecules such as carbon tetrachloride are discrete units which interact with adjacent molecules via intermolecular attractive forces. They can be comprised of two or more atoms, of the same or differing atoms.	
	(d) (iv)		
4	(a)	Covalent network solid	1
	(b)	$(SiC)_n$	1
	(c)	Covalent networks are similar in structure to ionic compounds forming a repeating three-dimensional structure. Covalent bonds are extremely strong, therefore bonds in solids with this arrangement require an extremely high amount of energy to break. This allows covalent network solids to withstand high degrees of mechanical stress.	1 1
	(d)	Non-conductive. As silicon carbide's bonding arrangement does not have any free electrons, the substance is unable to carry electricity. For a material that is to be worn on a the human body this is a desirable property.	1 1

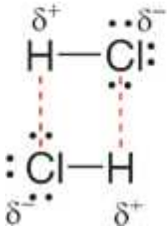
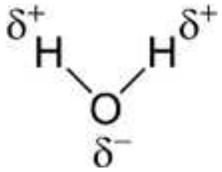
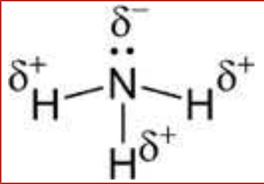
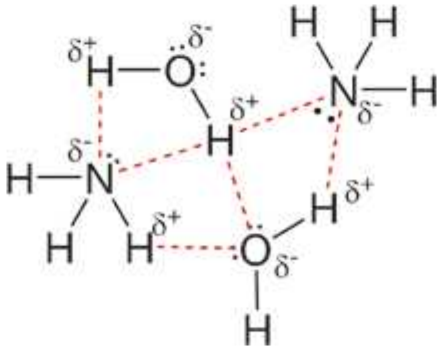
Question	Part	Author's response	Marks
5	(a)	$M = (6 \times M_C) + (8 \times M_H) + (6 \times M_O)$ $M = (6 \times 12.01) + (8 \times 1.008) + (6 \times 16)$ $M = 176.124 \text{ g mol}^{-1}$	1
		500 mg × 100 = 50 000 mg = 50 g $n = \frac{m}{M}$ $n = \frac{50}{176.124}$ $n = 2.84 \times 10^{-1}$ mol	1 1
		$N = n \times N_A$ $N = 2.84 \times 10^{-1} \times 6.02 \times 10^{23}$ $N = 1.71 \times 10^{23}$ molecules present in the sample	1 1
	(b)	m_c = % composition × 100 $m_c = 0.40 \times 100$ $m_c = 40.0$ g m_H = % composition × 100 $m_H = 0.0671 \times 100$ $m_H = 6.71$ g m_O = % composition × 100 $m_O = 0.5328 \times 100$ $m_O = 53.28$ g $n_C = \frac{m}{M}$ $n_C = \frac{40.0}{12.01}$ $n_C = 3.33$ mol $n_H = \frac{m}{M}$ $n_H = \frac{6.71}{1.008}$ $n_H = 6.66$ mol $n_O = \frac{m}{M}$ $n_O = \frac{53.28}{16.00}$ $n_O = 3.33$ mol $\therefore C_{6.16}H_{8.64}N_{1.23}$ $C = \frac{3.33}{3.33} \quad H = \frac{6.66}{3.33} \quad O = \frac{3.33}{3.33}$ Therefore, the empirical formula of the compound is CH_2O	

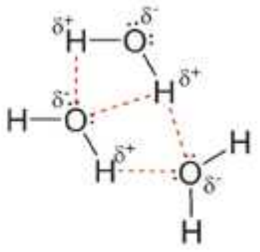
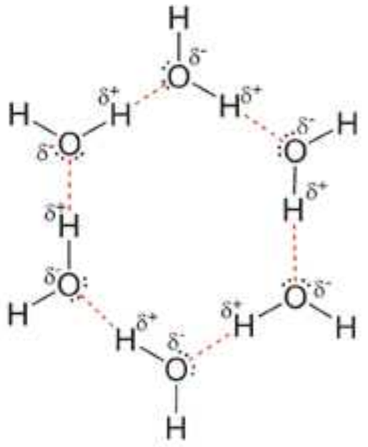
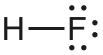
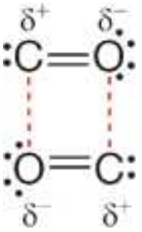
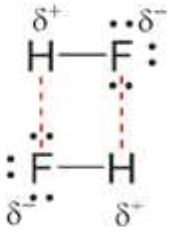
Topic 3: Molecules

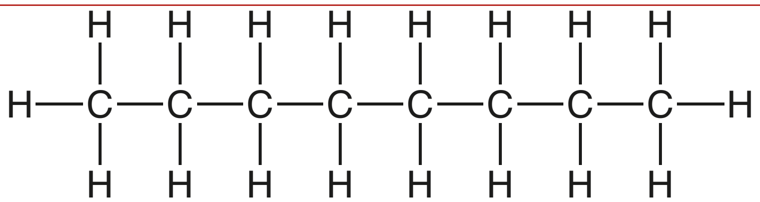
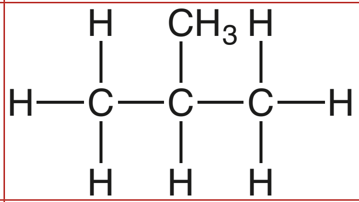
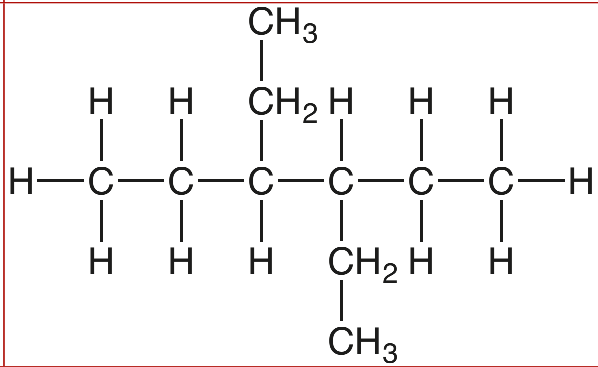
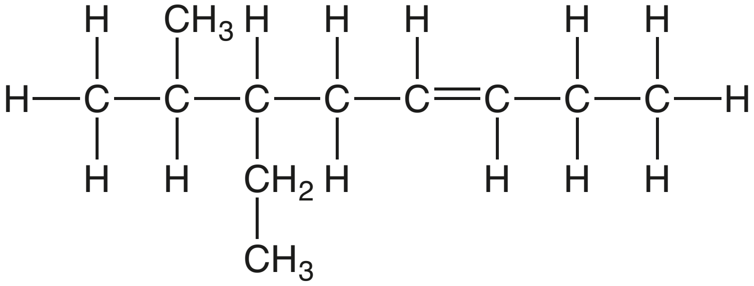
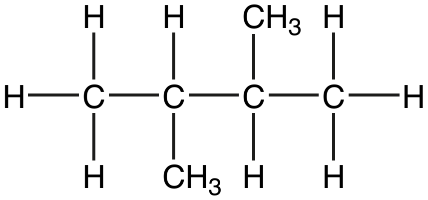
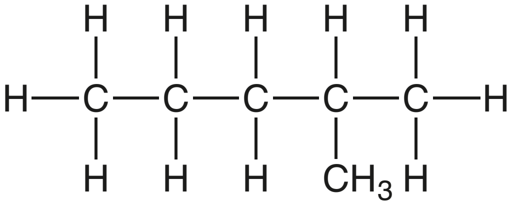
Question	Part	Author's response	Marks
1	(a)		1
	(b)		1
	(c)		1
	(d)		1
2	(a)	 tetrahedral	2
	(b)	 linear	2
	(c)	 v-shaped	2
	(d)	 trigonal pyramidal	2
3	(a)	Elements in period 3 of the periodic table have the ability to hold up to 18 electrons in their valence shell. These elements have the ability to share all of their valence electrons and thus acquire more than 8 electrons in their valence shell.	1 1
	(b)	Phosphorous is in period 3 of the periodic table and has 5 electrons in its third shell (valence). Its third shell can hold up to 18 electrons and thus, phosphorous can expand its octet and is able to share all 5 of its valence electrons with other atoms.	1 1

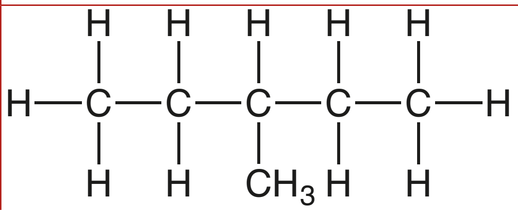
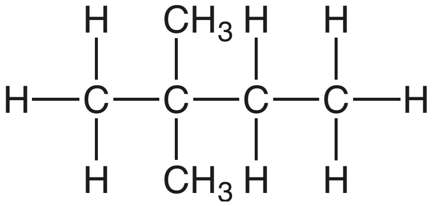
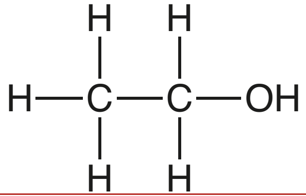
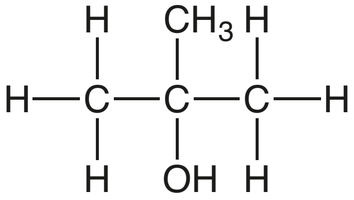
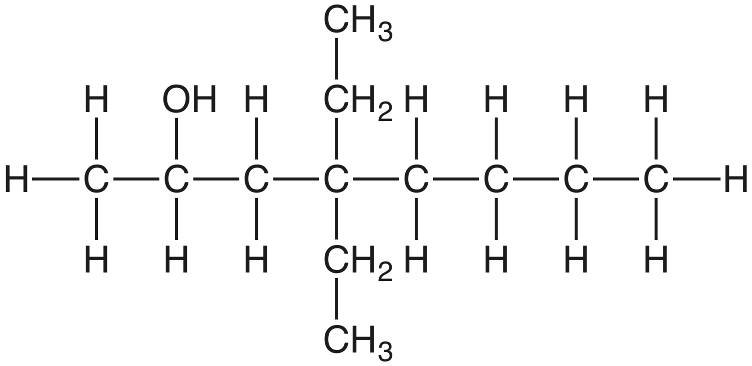
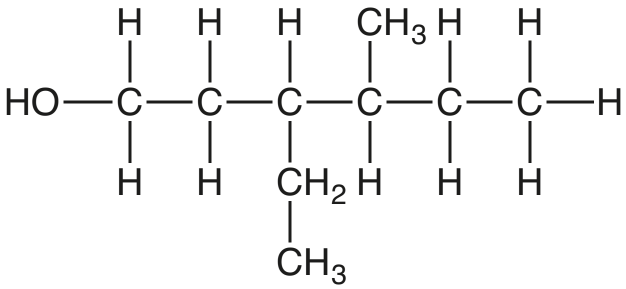
Question	Part	Author's response	Marks
4	(a)		2
	(b)		2
	(c)	$\overset{\delta^-}{O}=\overset{\delta^+}{C}=\overset{\delta^-}{O}$	2
	(d)	$H-\overset{\delta^+}{C}\equiv\overset{\delta^-}{N}$	2
5	(a)	 Polar Electric charge is unevenly distributed. Therefore, the molecule has distinct positive and negative poles.	1 1 1
	(b)	 Polar Electric charge is unevenly distributed. Therefore, the molecule has distinct positive and negative poles.	1 1 1
	(c)	 Non-polar Electric charge is evenly distributed. Therefore, there is no electric dipole.	1 1 1
	(d)	 Polar Electric charge is unevenly distributed. Therefore, the molecule has distinct positive and negative poles.	1 1

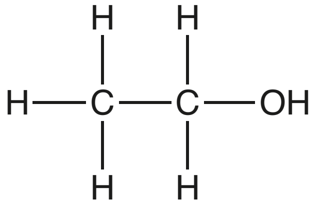
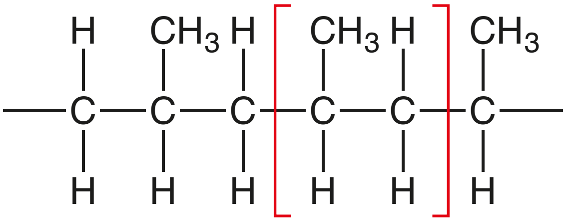
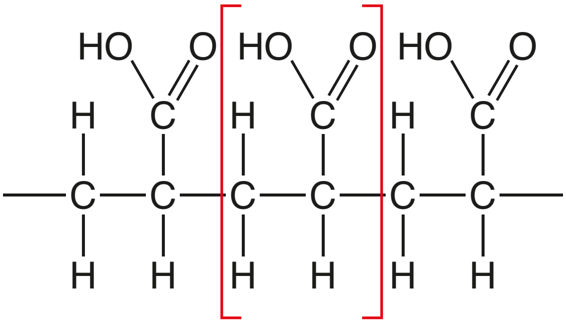
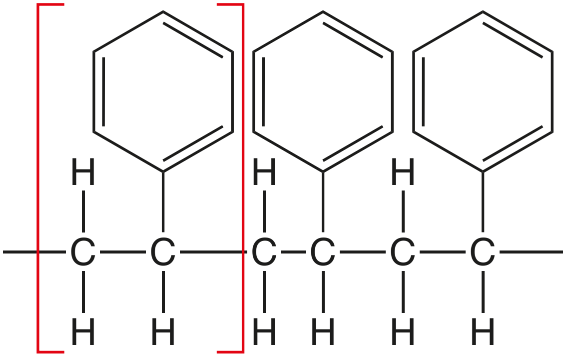
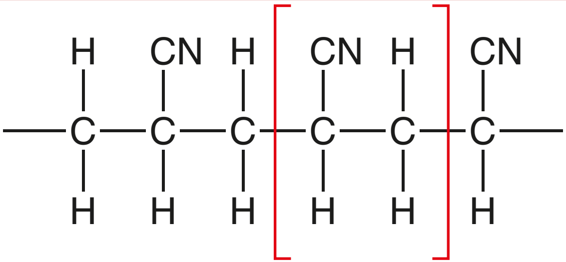
Question	Part	Author's response	Marks
6			1
			1
		Correct delta signs.	2
		Electric charge is evenly distributed within the molecule of carbon dioxide (CO_2). Therefore, there is no electric dipole.	1
		Electric charge is unevenly distributed within the molecule of sulfur dioxide (SO_2). Therefore, the molecule has distinct positive and negative poles.	1
7	(a)	KCl	1
		PCl_3	1
	(b)(i)	Potassium is a metal while chlorine is a non-metal, the bonding exhibited within this substance is ionic.	1
		Phosphorous and chlorine are both non-metals, the bonding exhibited within this substance is covalent.	1
	(ii)	Potassium has 1 valence electron and thus has a low electronegativity. Chlorine has 7 valence electrons and thus has high electronegativity. Chlorine will remove the valence electron from potassium.	1
		Phosphorous has 5 valence electrons. Chlorine has 7 valence electrons. Both have high electronegativities. Phosphorous and chlorine will share valence electrons.	1
	(c)		2
	(d)	Chlorine is more electronegative than phosphorous therefore, in each P-Cl bond the chlorine will have a partially negative charge. Phosphorous will have a partially positive charge.	1
		The centre of positive charge in the phosphorous trichloride molecule is positioned on the phosphorous atom.	1
		The centre of negative charge in the phosphorous trichloride molecule is positioned between the three chlorine atoms.	1
		This means that electric charge is unevenly distributed in the molecule which results in the molecule having distinct positive and negative poles; phosphorous trichloride is a polar molecule.	1
8	(a)	Intramolecular	1
	(b)	Intramolecular	1
	(c)	Intramolecular	1
	(d)	Intermolecular	1
9		Molecules of fluorine have 18 electrons orbiting their nuclei while molecules of hydrogen have 2 electrons orbiting their nuclei.	1
		Polarisation of electrons is greater in a molecule of fluorine which results in a stronger electric dipole.	1
		More thermal energy is required to separate molecules of fluorine, resulting in a higher boiling point.	1

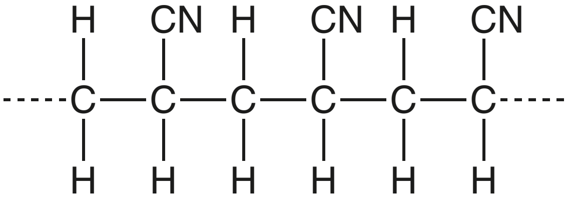
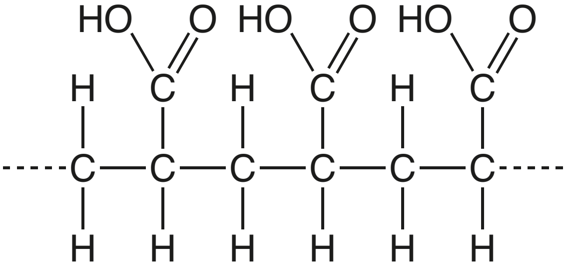
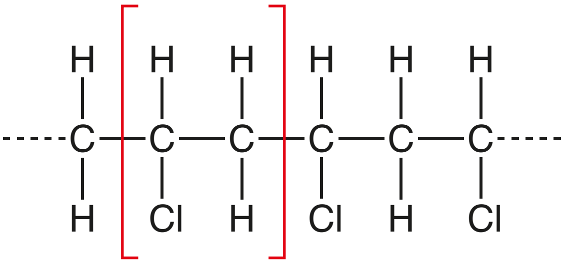
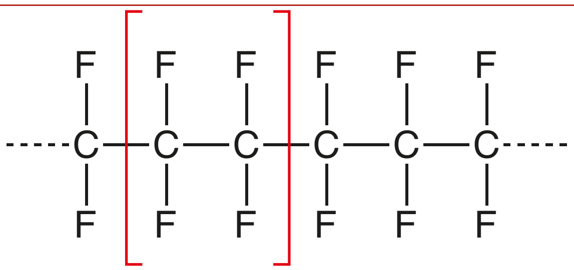
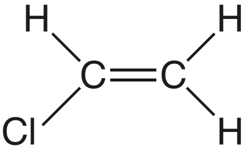
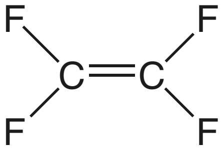
Question	Part	Author's response	Marks
10	(a)	Dipole-dipole interactions	1
	(b)	Dispersion forces	1
	(c)	 Hydrogen chloride is a polar molecule; secondary interactions acting between molecules of hydrogen chloride are dipole-dipole interactions. Fluorine is a non-polar molecule; secondary interactions acting between molecules of fluorine are dispersion forces. Dipole-dipole interactions are stronger; more thermal energy is required to separate molecules of hydrogen chloride. Hence, hydrogen chloride has a higher boiling point.	1 1 1 1
11	(a)	Dispersion forces	1
	(b)	Dipole-dipole interactions	1
	(c)	Dispersion forces	1
	(d)	Hydrogen bonding	1
12	(a)		2
	(b)		2
	(c)	 Water is a polar molecule and exhibits hydrogen bonding between adjacent molecules. Ammonia is a polar molecule and exhibits hydrogen bonding between adjacent molecules. Hydrogen bonding is exhibited between molecules of water and molecules of ammonia.	1 1 1 1

Question	Part	Author's response	Marks
12	(d)	 liquid (4°C)  solid (0°C)	2
		According to particle theory as thermal energy is removed from a substance the particles move closer together.	1
		From 4°C to 0°C water molecules move further apart due to the hydrogen bonds existing between molecules of water. The water molecules move into a regular array of molecules with the partially negative oxygen atom of each water molecule attracted to the partially positive hydrogen atom.	1
13	(a)	$2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(l)}$	1+1
	(b)	$C_3H_{8(g)} + 5O_{2(g))} \rightarrow 3CO_{2(g)} + 4H_2O_{(l)}$	1+1
	(c)	$2C_6H_{14(g)} + 18O_{2(g))} \rightarrow 12CO_{2(g)} + 14H_2O_{(l)}$	1+1
14	(a)	Incomplete combustion	1
	(b)	$2C_8H_{18} + 17O_2 \rightarrow 16CO + 18H_2O$	2
	(c)		2
		linear	1
	(d)	Hydrogen fluoride has one bonding electron pair (between the H and F atoms) as well as three non-bonding pairs of electrons.	1
		The four electron pairs are arranged into a shape which minimises repulsion around the central fluorine atom and the resulting shape is termed a linear molecule.	1
	(e)		1
			1
		Secondary interactions that exist between molecules of carbon monoxide are dipole-dipole interactions, while hydrogen bonds are present between adjacent molecules of hydrogen fluoride.	1
		Hydrogen bonding is a stronger form of secondary interactions and thus, more thermal energy is required to separate molecules of hydrogen fluoride.	1

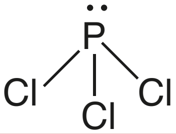
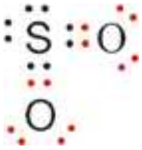
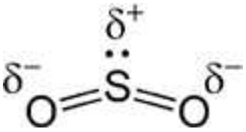
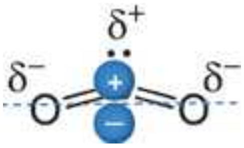
Question	Part	Author's response	Marks
15	(a)		2
	(b)		2
	(c)		2
	(d)		2
16	(a)	2-methylhexane	2
	(b)	ethene	2
	(c)	prop-3-yne	2
	(d)	2,2,4-trimethylheptane	2
17		 2,3-dimethylbutane	2+2
		 2-methylpentane	2+2

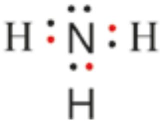
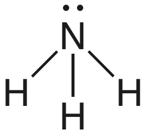
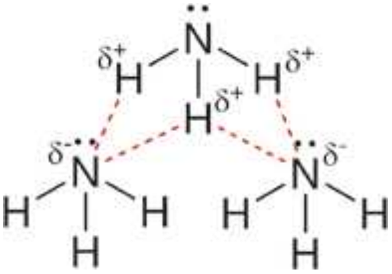
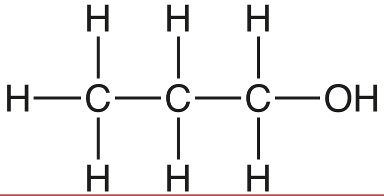
Question	Part	Author's response	Marks
17		 $\begin{array}{ccccccccccccc} & & H & & H & & H & & H & & H \\ & & \vert & & \vert & & \vert & & \vert & & \vert \\ H & - & C & - & C & - & C & - & C & - & C & - & H \\ & & \vert & & \vert & & \vert & & \vert & & \vert \\ & & H & & H & & CH_3 & & H & & H \end{array}$ 3-methylpentane	2+2
		 $\begin{array}{ccccccccccc} & & H & & CH_3 & & H & & H \\ & & \vert & & \vert & & \vert & & \vert \\ H & - & C & - & C & - & C & - & C & - & H \\ & & \vert & & \vert & & \vert & & \vert \\ & & H & & CH_3 & & H & & H \end{array}$ 2,2-dimethylbutane	2+2
18	(a)	 $\begin{array}{ccccccc} & & H & & H \\ & & \vert & & \vert \\ H & - & C & - & C & - & OH \\ & & \vert & & \vert \\ & & H & & H \end{array}$	2
	(b)	 $\begin{array}{ccccccccc} & & H & & CH_3 & & H \\ & & \vert & & \vert & & \vert \\ H & - & C & - & C & - & C & - & H \\ & & \vert & & \vert & & \vert \\ & & H & & OH & & H \end{array}$	2
	(c)	 $\begin{array}{ccccccccccccccccccc} & & & & & & & & CH_3 \\ & & & & & & & & \vert \\ & & H & & OH & & H & & CH_2 & & H & & H & & H & & H \\ & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert \\ H & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & C & - & H \\ & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert \\ & & H & & H & & H & & CH_2 & & H & & H & & H & & H \\ & & & & & & & & \vert \\ & & & & & & & & CH_3 \end{array}$	2
	(d)	 $\begin{array}{ccccccccccccccc} & & H & & H & & H & & CH_3 & & H & & H \\ & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert \\ HO & - & C & - & C & - & C & - & C & - & C & - & C & - & H \\ & & \vert & & \vert & & \vert & & \vert & & \vert & & \vert \\ & & H & & H & & CH_2 & & H & & H & & H \\ & & & & & & \vert \\ & & & & & & CH_3 \end{array}$	2
19	(a)	butan-1-ol	2
	(b)	2-methylbutan-1-ol	2
	(c)	2-methylhexan-3-ol	2
	(d)	3,7-dimethyloctan-3-ol	2

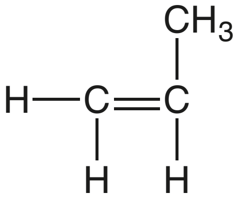
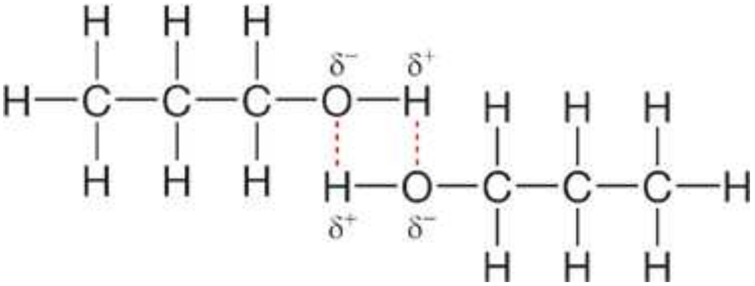
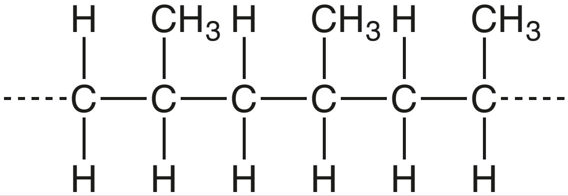
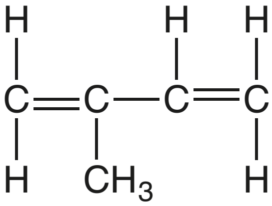
Question	Part	Author's response	Marks
20	(a)		2
	(b) i)	Ethanol has a small molecular structure. Ethyl butanoate has a much larger molecular structure. Ethyl butanoate has more 'sites' for the polarisation of electrons to occur and thus more dipole moments. Polarisation of electrons is greater in a molecule of ethyl butanoate and can occur more frequently which results in a stronger electric dipole. More thermal energy is required to overcome the stronger secondary forces of attraction between ethyl butanoate molecules and hence, the boiling point of ethyl butanoate is higher than that of ethanol.	1 1 1 1
	(ii)	Ethanol and ethyl butanoate have two distinct boiling points. Distillation is used to separate these two substances at their distinct boiling points	1 1
21	(a)		2
	(b)		2
	(c)		2
	(d)		2

Question	Part	Author's response	Marks
22	(a)		2
	(b)		2
23	(a)(i)		2
	(ii)		2
	(b)(i)		2
	(ii)		2
24		Cyanoacrylate exhibits secondary interactions between polymer chains.	1
		Epoxy resin exhibits cross-linking between polymer chains which are stronger covalent bonds.	1
		A very large amount of heat energy is required to break-down the epoxy resin structure and separate the polymer chains.	1
		Rubber will retain its shape at high temperatures as the cross-link covalent bonds ensure the chains cannot slip over/past each other.	1

Summary Test 3: Molecules

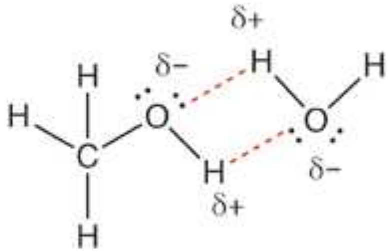
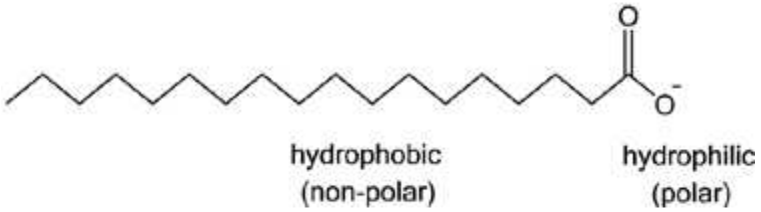
Question	Part	Author's response	Marks
1	(a)	silicon chloride	1
	(b)	Molecules Weak Ions Strong	1 1 1 1
	(c)		2
2	(a)(i)	2	1
	(ii)	2	1
	(b)		1
	(c)	Elements in period 3 of the periodic table are able to hold up to 18 electrons in their valence shell. An element has expanded its octet when it acquires more than 8 electrons in its valence shell. Sulfur is in period 3 of the periodic table and has 6 electrons in its third shell (valence). Its third shell can hold up to 18 electrons and thus, phosphorous can expand its octet and is able to share all 6 of its valence electrons with other atoms.	1 1 1 1
	(d)	 Correct delta signs	2 1
	(e)	Oxygen is more electronegative than sulfur. Due to the unequal distribution of charge, in each S-O bond the oxygen atom will have a partially negative charge and the sulfur atom will have a partially positive charge.	1 1
	(f)	 Sulfur dioxide is a polar molecule. Electric charge is unevenly distributed within the molecule of sulfur dioxide (SO_2). Therefore, the molecule has distinct positive and negative poles.	1 2
	(g)	Dipole-dipole interactions	1

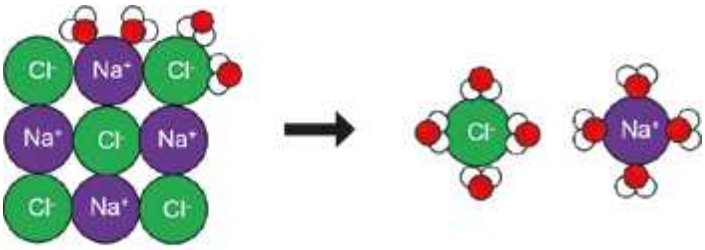
Question	Part	Author's response	Marks
3	(a)	$:N:::N:$	1
	(b)	$:N\equiv N:$ Linear	2 1
	(c)	Nitrogen gas exhibits a strong N-N triple covalent bond within its molecular structure. A large amount of energy is needed to break this bond hence rendering it a relatively unreactive gas.	1 1
	(d)		1
	(e)	 Trigonal pyramidal	2 1
	(f)	Ammonia has three bonding electron pairs (between the N and H atoms) as well as one non-bonding pair of electrons. The four electron pairs are arranged into a shape which minimises repulsion around the central nitrogen atom and the resulting shape is a trigonal pyramidal molecule.	1 1
	(g)	 Nitrogen is a non-polar molecule as electric charge is evenly distributed. Ammonia is a polar molecule as electric charge is unevenly distributed. The secondary interactions that exist between molecules of ammonia are hydrogen bonds which are much stronger than the dispersion forces exhibited between molecules of nitrogen. A larger amount of heat energy is required to separate molecules of ammonia. Thus, the boiling point is higher.	1 1 1 1 1
4	(a)	 $H-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-\overset{H}{\underset{H}{C}}-OH$	2
	(b)	Acetone is a relatively non-polar molecule and the secondary interactions that exist between molecules of acetone are weak dispersion forces. A small amount of heat energy is required to separate molecules of acetone. Thus, acetone has a low boiling point.	1 1

Question	Part	Author's response	Marks
5	(a)		2
	(b)		1
		Propene is a non-polar molecule as electric charge is evenly distributed across the molecule.	1
		Propan-1-ol is a relatively small hydrocarbon and has a polar hydroxyl group. Therefore, electric charge is unevenly distributed within the molecule and it has distinct positive and negative poles.	1
		The secondary interactions that exist between molecules of propan-1-ol are hydrogen bonds which are much stronger than the dispersion forces exhibited between molecules of propene.	1
		A larger amount of heat energy is required to separate molecules of propan-1-ol. Thus, the boiling point is higher.	1
	(c)		2
	(d)	Both propene and polypropene are non-polar molecules due to electric charge being evenly distributed across both molecules. Therefore, both molecules exhibit dispersion forces as the secondary forces of attraction acting between them.	1
		Polypropene has more 'sites' for the polarisation of electrons to occur and thus more dipole moments.	1
		Polarisation of electrons is greater in a molecule of polypropene and can occur more frequently which results in a stronger electric dipole.	1
		More thermal energy is required to overcome the stronger secondary forces of attraction between polypropene molecules and hence, the boiling point of polypropene is much higher than that of propene.	1
6	(a)		2
	(b)	After vulcanisation rubber exhibits cross-linking between polymer chains. These are strong covalent bonds between carbon and sulfur atoms.	1
		A very large amount of heat energy is required to break-down this structure of rubber and separate the polymer chains.	1
		Rubber will retain its shape at high temperatures as the cross-link covalent bonds ensure the chains cannot slip over/past each other.	1

Topic 4: Mixtures and Solutions

Question	Part	Author's response	Marks
1	(a)	The molten zinc (solute) is the substance that is dissolved/minority component of the solution. The molten copper (solvent) is the substance in which the solute is dissolved in/majority component of the solution.	1 1
2	(a)	Miscibility describes liquids which can be mixed together to form a single liquid. Gasoline is formed from miscible components.	1 1
	(b)	Immiscible solution.	1
3	(a)	Hydrogen bonding. Presence of the polar N-H group.	1 1
	(b)	Water contains a polar hydroxyl group. Also capable of hydrogen bonding.	1 1
	(c) i)	Propan-1-ol is polar and methane is non-polar. As water is also polar, the intermolecular interactions can be overcome allowing the particles of both to separate and mix together. This forms new hydrogen bonds between the propan-1-ol and water molecules and resulting in a miscible solution.	1 1 1
	ii)	Hydrochloric acid is polar, iodine pentachloride is non-polar. This will have the overall result that hydrochloric acid is miscible in water. Iodine pentachloride has relatively weak intermolecular attractions (dispersion forces) between molecules that cannot overcome the stronger hydrogen bonds that hold the water molecules together, resulting in an immiscible solution.	1 1 1
4	(a)	Water is polar. Hexane is non-polar with its molecules attracted to one another by dispersion forces. As the substance is observed to be miscible in hexane, its intermolecular attractions must match those of hexane, resulting in the fact that it is most likely to be non-polar in nature.	1 1 1
	(b)	Solvents and solutes of similar (like) size and polarity mix readily.	1
5	(a)	Immiscible. Extensive network of intermolecular forces between benzene molecules. Methanol molecules cannot separate and surround molecules of benzene.	1 1 1
	(b)	Miscible. Both molecules are small and polar. When mixed, these interactions are separated and replaced with new interactions (hydrogen bonds) between the propan-1-ol and water molecules resulting in a miscible solution.	1 1 1
	(c)	Miscible. Molecules are non-polar. As these interactions can be overcome and therefore be separated and replaced with intermolecular interactions between molecules of octane and cyclohexane, resulting in a homogenous solution.	1 1 1

Question	Part	Author's response	Marks
6	(a)		2
	(b)	Methanol has non-polar component. Dispersion forces form between non-polar regions of methanol and hexane allowing the materials to mix.	1 1
	(c)	No/immiscible. The non-polar region of heptan-1-ol is capable only of weak dispersion forces that would not be able to overcome the stronger hydrogen bonds that hold water molecules together.	1 1
	(d)	Molecules must be small and polar. Most organic hydrocarbons are large and non-polar.	1 1
7	(a)	An immiscible solution would form. Extensive network of intermolecular forces between vitamin K molecules. Water molecules cannot separate and surround molecules of vitamin K.	1 1 1
	(b)	Fat soluble	1
8	(a)		2
	(b)	The hydrophilic head attracts water. The hydrophobic tail does not attract water.	1 1
	(c)	Emulsifier (detergent) has polar and non-polar regions. Polar (hydrophilic) region attracts water. Non-polar region attracts non-polar substance.	1 1 1
9	(a)	Emulsifier (diglyceride) has polar and non-polar regions. Polar (hydrophilic) region attracts water molecules. Non-polar (hydrophobic) backbone attracts non-polar components of the ice-cream.	1
	(b)	An emulsion is a mixture of two immiscible liquids. An emulsifier assists an emulsion in becoming more stable by allowing the oil droplets to stay dispersed in the water for a longer time.	1 1

<table>
<tr><th>Question</th><th>Part</th><th>Author's response</th><th>Marks</th></tr>
<tr><td rowspan="9">10</td><td rowspan="2">(a)</td><td>Salts are ionic compounds that are formed in acid-base reactions.</td><td>1</td></tr>
<tr><td>Calcium (Ca^{2+}) cations attract chloride (Cl^-) anions forming a neutral (ionic) compound in a 2:1 ratio.</td><td>1</td></tr>
<tr><td rowspan="3">(b)</td><td>Hydration occurs.</td><td>1</td></tr>
<tr><td>Attraction of solvent molecules (water) to the ions of the solute (calcium chloride).</td><td>1</td></tr>
<tr><td>Ions dissolve and spread out, being surrounded by solvent molecules.</td><td>1</td></tr>
<tr><td>(c)</td><td>
</td><td>4</td></tr>
<tr><td>(d)</td><td>Ion-ion attraction is stronger than the water-ion attraction.</td><td>1</td></tr>
<tr><td colspan="3"></td></tr>
<tr><td colspan="3"></td></tr>
<tr><td>11</td><td></td><td>
<table>
<tr><th>Salt</th><th>Soluble in water</th><th>Insoluble in water</th></tr>
<tr><td>iron(III) nitrate</td><td>✓</td><td></td></tr>
<tr><td>copper(II) sulfate</td><td>✓</td><td></td></tr>
<tr><td>mercury(II) sulfide</td><td></td><td>✓</td></tr>
<tr><td>zinc chloride</td><td>✓</td><td></td></tr>
<tr><td>ammonium phosphate</td><td>✓</td><td></td></tr>
<tr><td>lead(II) acetate</td><td>✓ (moderate)</td><td></td></tr>
</table>
</td><td>6</td></tr>
<tr><td rowspan="11">12</td><td>(a)</td><td>No dissociation</td><td>2</td></tr>
<tr><td>(b)</td><td>$NaCH_3COO_{(s)} + aq \rightarrow Na^+_{(aq)} + CH_3COO^-_{(aq)}$</td><td>2</td></tr>
<tr><td>(c)</td><td>$(NH_4)_2SO_{4(s)} + aq \rightarrow 2NH^+_{4(aq)} + SO^{2-}_{4(aq)}$</td><td>2</td></tr>
<tr><td>(d)</td><td>No dissociation</td><td>2</td></tr>
<tr><td>(e)</td><td>$K_2CO_{3(s)} + aq \rightarrow 2K^+_{(aq)} + CO^{2-}_{3(aq)}$</td><td>2</td></tr>
<tr><td>(f)</td><td>$NaOH_{(s)} + aq \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$</td><td>2</td></tr>
<tr><td>(g)</td><td>$FeSO_{4(s)} + aq \rightarrow Fe^{2+}_{(aq)} + SO^{2-}_{4(aq)}$</td><td>2</td></tr>
<tr><td>(h)</td><td>No dissociation</td><td>2</td></tr>
<tr><td>(i)</td><td>No dissociation</td><td>2</td></tr>
<tr><td>(j)</td><td>$FeCl_{3(s)} + aq \rightarrow Fe^{3+}_{(aq)} + Cl^-_{(aq)}$</td><td>2</td></tr>
<tr><td>(k)</td><td>$Ca(NO_3)_{2(s)} + aq \rightarrow Ca^{2+}_{(aq)} + 2NO^-_{3(aq)}$</td><td>2</td></tr>
<tr><td rowspan="7">13</td><td>(a)</td><td>$NH_4NO_{3(s)} + aq \rightarrow NH_{4(aq)} + NO^-_{3(aq)}$</td><td>2</td></tr>
<tr><td rowspan="2">(b)</td><td>Soluble salt.</td><td>1</td></tr>
<tr><td>Releases nitrates into soil solution.</td><td>1</td></tr>
<tr><td rowspan="2">(c)</td><td>Dissociation produces free nitrate anions and ammonium cations.</td><td>1</td></tr>
<tr><td>Ion-dipole interactions are formed between anions/cations and molecules of water.</td><td>1</td></tr>
<tr><td rowspan="2">(d)</td><td>Soluble in water.</td><td>1</td></tr>
<tr><td>Urea will form hydrogen bonds between the carbonyl group and adjacent water molecules via the dipoles present on both neutral molecules.</td><td>1</td></tr>
</table>

Question	Part	Author's response	Marks
14	(a) i)	Potassium nitrate.	1
	ii)	$Pb^{2+}_{(aq)} + 2I^{-}_{(aq)} \rightarrow PbI_{2(s)}$	2
	iii)	Potassium and nitrate ions.	1
	iv)	Precipitation.	1
	(b)	They do not undergo chemical change in the reaction.	1
	(c)	Both potassium nitrate and potassium iodide dissociate when placed in water. The ions then bond with the respective anion or cation from the other reactant.	1 1
15	(a)	Sodium sulfate	1
	(b)	$2NaOH_{(aq)} + NiSO_{4(aq)} \rightarrow Ni(OH)_{2(s)} + 2Na^{+}_{(aq)} + SO^{2-}_{4(aq)}$	2
	(c)	Sodium and sulfate ions	1
	(d)	$Ni^{2+}_{(aq)} + 2OH^{-}_{(aq)} \rightarrow Ni(OH)_{2(s)}$	2
	(e)	The sulfate and hydroxide ions "replace" each other's cation partner.	1
16	(a) i)	$Fe(OH)_{2(s)}$, $Na^{+}_{(aq)}$ and $NO^{-}_{3(aq)}$	1
	ii)	$Fe(NO_3)_{3(aq)} + 3NaOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3Na^{+}_{(aq)} +3NO^{-}_{3(aq)}$	2
	iii)	$Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)} \rightarrow Fe(OH)_{3(s)}$	2
	(b) i)	$BaSO_{4(s)}$, $Al^{3+}_{(aq)}$ and $Cl^{-}_{(aq)}$	1
	ii)	$Al_2(SO_4)_{3(aq)} + 3BaCl_{2(aq)} \rightarrow 3BaSO_{4(s)} + 2Al^{3+}_{(aq)} + 6Cl^{-}_{(aq)}$	2
	iii)	$Ba^{2+}_{(aq)} + SO^{2+}_{4(aq)} \rightarrow BaSO_{4(s)}$	2
	(c) i)	$HNO_{3(aq)}$ and $ZnI_{2(aq)}$	1
	ii)	$2HI_{(aq)} + Zn(NO_3)_{2(aq)} \rightarrow 2H^{+}_{(aq)} + 2NO^{-}_{3(aq)} + Zn^{2+}_{(aq)} + 2I^{-}_{(aq)}$	2
	iii)	Not applicable as no precipitation occurs	2
17	(a)	$2Al_{(s)} + 3O_{2(g)} \rightarrow Al_2O_{3(s)}$	1
	(b)	$CaI_{2(aq)} + Hg(NO_3)_{2(aq)} \rightarrow Ca(NO_3)_{2(aq)} + HgI_{2(s)}$	1
	(c)	$Mg_{(s)} + 2CH_3COOH_{(aq)} \rightarrow Mg(CH_3COO)_{2(aq)} + H_{2(g)}$	1
	(d)	$CoSO_{3(s)} \rightarrow CoO_{(s)} + SO_{2(g)}$	1
	(e)	$Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 2Fe_{(s)} + 3CO_{(g)}$	1
	(f)	$2Cu_{(s)} + S_{(s)} \rightarrow Cu_2S_{(s)}$	1
	(g)	$2AgNO_{3(aq)} + K_2CrO_{4(s)} \rightarrow Ag_2CrO_{4(s)} + 2KNO_{3(aq)}$	1
	(h)	$Sr(OH)_{2(aq)} + H_2SO_{4(aq)} \rightarrow SrSO_{4(s)} + 2H_2O_{(l)}$	1
	(i)	$2Pb(NO_3)_{2(aq)} \rightarrow 2PbO_{(s)} + 4NO_{2(g)} + O_{2(g)}$	1
	(j)	$2NH_{3(g)} + H_2SO_{4(aq)} \rightarrow (NH_4)2SO_{4(aq)}$	1
18	(a)	$2AgNO_{3(aq)} + MgCl_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2AgCl_{(s)}$	2
	(b)	$3Zn_{(s)} + 2CoCl_{3(aq)} \rightarrow 3ZnCl_{2(aq)} + 2Co_{(s)}$	2
	(c)	$Fe_2O_{3(s)} + 6HCl_{(aq)} \rightarrow 2FeCl_{3(aq)} + 3H_2O_{(l)}$	2
	(d)	$6NaOH_{(aq)} + Al_2(SO_4)_{3(aq)} \rightarrow 3Na_2SO_{4(aq)} + 2Al(OH)_{3(s)}$	2
	(e)	$Pb(NO_3)_{2(aq)} + 2HCl_{(aq)} \rightarrow PbCl_{2(s)} + 2HNO_{3(aq)}$	2
	(f)	$CuSO_{4(aq)} + Cd(NO_3)_{2(aq)} \rightarrow CdSO_{4(aq)} + Cu(NO_3)_{2(aq)}$	2

Question	Part	Author's response	Marks
19	(a)	$n_{(Ag_2NO_3)} = \frac{0.9}{169.91}$ $n_{(Ag_2NO_3)} = 5.3 \times 10^{-3}$ mol	
	(b)	$c_{(Ag_2NO_3)} = \frac{5.3 \times 10^{-3}}{0.250}$ $c_{(Ag_2NO_3)} = 0.02$ mol L^{-1}	1 1
	(c)	$c_{(Ag_2NO_3)} = 0.02 \times 169.91$ $c_{(Ag_2NO_3)} = 3.6$ g L^{-1}	
	(d)	Concentration refers to the number of particles within a given volume of solution. A concentrated solution has a greater number of particles than a dilute solution of the same volume of solution.	1 1
20	(a)	$c_{(NaCl)} = \frac{m}{V}$ $c_{(NaCl)} = \frac{16.0}{0.500}$ $c_{(NaCl)} = 32$ g L^{-1}	1 1
	(b)	$c_{(NaCl)} = \frac{35{,}000}{1000}$ $c_{(NaCl)} = 0.60$ mol L^{-1}	1 1
	(c)	$c_{(NaCl)} = 0.60$ mol L^{-1} × 58.44 × 1000 $c_{(NaCl)} = 32{,}000$ ppm 32,000 ppm $\neq$ 35,000 ppm The student's solution did not successfully mimic the concentration of sea water.	1 1 1
21	(a)	$n_{(Al)} = \frac{m}{M}$ $n_{(Al)} = \frac{5.4}{26.98}$ $n_{(Al)} = 0.2$ mol Using mole ratio: $\frac{n_{(HCl)}}{n_{(Al)}} = \frac{6}{2}$ $n_{(HCl)} = \frac{6}{2} \times 0.2$ $n_{(HCl)} = 0.6$ mol $V_{(HCl)} = \frac{0.6}{2}$ $V_{(HCl)} = 0.3$ L (300 mL)	1 1 1 1 1 1
	(b)	Using mole ratio: $\frac{n_{(H_2)}}{n_{(HCl)}} = \frac{3}{6}$ $n_{(H_2)} = \frac{3}{6} \times 0.6$ $m_{(H_2)} = 0.3$ mol $m_{(H_2)} = n.M$ $m_{(H_2)} = 0.3 \times 2.016$ $m_{(H_2)} = 0.6$ g	1 1 1 1 1

Question	Part	Author's response	Marks
22	(a)	$g\ L^{-1} = mol\ L^{-1} \times M$ $g\ L^{-1} = 4.5 \times 10^{-3} \times 180.156$ $= 0.81\ g\ L^{-1}$	1 + 1 1
	(b)	$ppm = g\ L^{-1} \times 10^3$ $ppm = 0.81 \times 10^3$ $= 811\ ppm$	1 1
23	(a)	$g\ L^{-1} = \frac{mg\ L^{-1}}{10^3}$ $g\ L^{-1} = \frac{4}{10^3}$ $= 4 \times 10^{-3} g\ L^{-1}$ %w/v = $\frac{g\ L^{-1}}{10}$ %w/v = $\frac{4 \times 10^3}{10}$ $= 4 \times 10^{-4}$ %w/V	1 1 1 1
	(b)	$ppb = g\ L^{-1} \times 10^6$ $ppb = 4 \times 10^6$ $= 4 \times 10^6\ ppb$	1 1
24	(a) i)	$c_iV_i = c_fV_f$ $V_i = \frac{c_f.V_f}{c_i}$ $V_i = \frac{0.2 \times 0.5}{0.25}$ $V_i = 0.4$ L (400 mL)	1 1
	ii)	$V_i = \frac{0.2 \times 0.5}{1.0}$ $V_i = 0.1$ L (100 mL)	1 1
	iii)	$V_i = \frac{0.2 \times 0.5}{2.5}$ $V_i = 0.04$ L (40 mL)	1 1
	(b)	Using a 500 mL volumetric flask and funnel, add 400 mL of 0.25 mol L^{-1} stock solution NaOH. Fill to the calibration mark with distilled water, using a teat pipette final few mLs to ensure a good degree of accuracy.	1 1
	(c) i)	$V_i = \frac{0.2 \times 2}{18}$ $V_i = 0.022$ L (22 mL)	1 1
	ii)	$V_i = \frac{0.2 \times 4.25}{15.8}$ $V_i = 0.054$ L (54 mL)	1 1

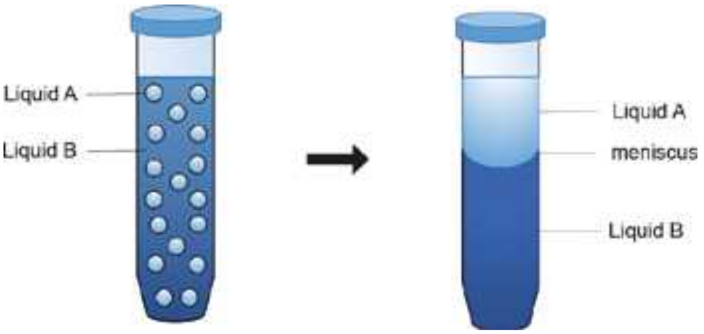
Question	Part	Author's response	Marks
25		$n_{(Na_2CO_3)} = \frac{m}{M}$ $n_{(Na_2CO_3)} = \frac{21.2}{105.99}$ $n_{(Na_2CO_3)} = 0.2$ mol $c_{(Na_2CO_3)} = \frac{n}{v}$ $c_{(Na_2CO_3)} = \frac{0.2}{0.25}$ $c_{(Na_2CO_3)} = 0.8$ mol L^{-1}	1 1 1 1
		Use a clean watch glass and mass balance to accurately measure 21.2 g of the standard (sodium carbonate).	1
		Using a spatula and funnel, transfer the solid to a 250 mL volumetric flask. Use a small amount of distilled water to rinse the solid from the watch glass into the flask to ensure complete transfer of the standard.	1
		Continue to add distilled water, and swirl to ensure the standard is dissolved. When approaching the calibration mark, use a teat pipette to deliver the final volume of distilled water, taking care to ensure the bottom of the meniscus is resting on the calibration mark.	1
		Stopper and invert the solution to ensure a homogenous solution.	1
26	(a)	$Mg_{(s)} + CuSO_{4(aq)} \longrightarrow MgSO4_{(aq)} + Cu_{(s)}$	2
	(b)	$n_{(CuSO_4)} = \frac{m}{Mm}$ $n_{(CuSO_4)} = \frac{17}{159.61}$ $n_{(CuSO_4)} = 0.11$ mol	1 1
	(c)	Using mole ratio: $\frac{n_{(Cu)}}{n_{(CuSO_4)}} = \frac{1}{1}$ $n_{(Cu)} = 0.11$ mol	1 1
	(d)	$m_{(Cu)} = n \times M$ $m_{(Cu)} = 0.11 \times 63.55$ $m_{(Cu)} = 6.77$ g	1 1
27	(a)	Aluminium chloride	1
	(b)	$2Al_{(s)} + 3ZnCl_{2(s)} \longrightarrow 3Zn_{(s)} + 2AlCl_{3(aq)}$	2
	(c) i)	$n_{(AlCl_3)} = \frac{14.7}{133.33}$ $n_{(AlCl_3)} = 0.11$ mol	1 1
	ii)	Using mole ratio: $\frac{n_{(ZnCl_2)}}{n_{(AlCl_3)}} = \frac{3}{2}$ $n_{(ZnCl_2)} = \frac{3}{2} \times 0.11$ $n_{(ZnCl_2)} = 0.155$ mol	1 1
	iii)	$m_{(ZnCl_2)} = n \times M$ $m_{(ZnCl_2)} = 0.166 \times 136.315$ $m_{(ZnCl_2)} = 22.5$ g	1+1 1

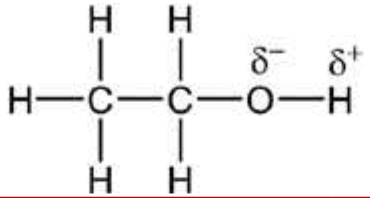
Question	Part	Author's response	Marks
28		$2KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$ $n_{(PbI_2)} = \frac{1.67}{461.01}$ $n_{(PbI_2)} = 3.6 \times 10^{-3}$ mol Using mole ratio: $\frac{n_{(KI)}}{n_{(PbI_2)}} = \frac{2}{1}$ $n_{(KI)} = \frac{2}{1} \times 3.62 \times 10-3$ mol $n_{(KI)} = 7.24 \times 10^{-3}$ mol $m_{(KI)} = 7.24 \times 10^{-3} \times 166.003$ $m_{(KI)} = 1.20$ g	2 1 1 1 1 1+1 1
29		$n_{(AgCl)} = \frac{0.115}{143.32}$ $n_{(AgCl)} = 8.02 \times 10^{-4}$ mol Using mole ratio: $\frac{n_{(NaCl)}}{n_{(AgCl)}} = \frac{1}{1}$ $n_{(NaCl)} = 8.02 \times 10^{-4}$ mol $m_{(NaCl)} = n \times M$ $m_{(NaCl)} = 8.02 \times 10^{-4} \times 58.44$ $m_{(NaCl)} = 0.0469$ g	1 1 1+1 1
30	(a) i)	Exothermic	1
	ii)	$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$ $\Delta H = -890$ kJ mol^{-1}	4
	(b) i)	Exothermic	1
	ii)	$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$ $\Delta H = -286$ kJ mol^{-1}	4
	(c) i)	Endothermic	1
	ii)	$6H_2O_{(l)} + 6CO_{2(g)} \rightarrow C_6H_{12}O_{6(l)} + 6O_{2(g)}$ $\Delta H = +2802$ kJ mol^{-1}	4
31	(a)	Endothermic	1
	(b)	Absorbed	1
	(c)	Energy of reactants < energy of the products	1
	(d)	For a chemical reaction to proceed, successful or productive collisions must take place. For a collision to be productive, particles must possess sufficient energy to overcome the activation energy of the reaction, thus energy is imperative to the progression of chemical reactions.	1 1
	(e)	Total energy of system (cold pack) before reaction equals the total energy after reaction. Energy changes forms during the the reaction when the membrane is broken, but no energy is created and none is destroyed.	1 1
	(f)	Hydration energy = lattice energy - ΔH = 705 – 17 = 688 kJ mol^{-1}	1
	(g)	Lattice energy is the energy required to dissociate one mole of the ammonium chloride into free ions. Hydration energy refers to the energy released as molecules of water adhere to chloride anions and ammonium cations. As this amount is less than the lattice energy, the dissociation of ammonium chloride is endothermic.	1 1 1

Question	Part	Author's response	Marks
32	(a)	NaCl: $\Delta H_{(solution)} = 779 - 774 = +5\ kJ\ mol^{-1}$	1
		KCl: $\Delta H_{(solution)} = 701 - 685 = +16$ kJ mol-1	1
		LiF: $\Delta H_{(solution)} = 1032 - 1005 = +27\ kJ\ mol^{-1}$	1
		NaOH: $\Delta H_{(solution)} = 737 - 799 = -62\ kJ\ mol^{-1}$	1
		AgCl: $\Delta H_{(solution)} = 916 - 851 = +65\ kJ\ mol^{-1}$	1
	(b)	A negative enthalpy of solution like the one demonstrated when sodium iodide dissociates in water indicates that the hydration energy is greater than that of the lattice energy. Or in other words more energy is released when water surrounds the ions Than what is absorbed to pull apart the ions resulting in its dissociation.	1 1 1
33	(a) i)	$Q = m(water) \times \Delta T \times c_p(water)$ $Q = 25.5 \times 8.5 \times 4.18$ $Q = 906$ J	1 1 1
	ii)	$Q = m(water) \times \Delta T \times c_p(water)$ $Q = 75 \times 72.5 \times 4.18$ $Q = 22\ 729$ J $Q = 22.7$ kJ	1 1
	(b)	Specific heat capacity is a constant (for water this is 4.18 $J.g^{-1}.°C^{-1}$). It refers to the amount of energy (in joules) required to increase the temperature of 1 gram of water by 1°C.	1 1
34		$n_{(Ce)} = \frac{7.0}{132.9}$ $n_{(Ce)} = 5.3 \times 10^{-3}mol$ $\Delta H = \frac{m_{water} \times \Delta T_{water} \times c_{p\ (water)}}{1000 \times n_{substance}}$ $\Delta H = \frac{250 \times 1.02 \times 4.18}{1000 \times (5.3 \times 10^{-2})}$ $\Delta H = +20.2\ kJ\ mol^{-1}$	1 1 1 1
35		$n_{(NH_4NO_3)} = \frac{8.0}{80.043}$ $n_{(NH_4NO_3)} = 9.9 \times 10^{-2}mol$ $\Delta H = \frac{m_{water} \times \Delta T_{water} \times c_{p\ (water)}}{1000 \times n_{substance}}$ $\Delta H = \frac{1000 \times 0.61 \times 4.18}{1000 \times (9.9 \times 10^{-2})}$ $\Delta H = +25.5\ kJ\ mol^{-1}$	1 1 1 1

Question	Part	Author's response	Marks
36	(a)	$n_{(Mg)} = \frac{0.1277}{24.3}$ $n_{(Mg)} = 5.26 \times 10^{-3}$mol $\Delta H = \frac{m_{water} \times \Delta T_{water} \times c_{p\ (water)}}{1000 \times n_{substance}}$ $\Delta H = \frac{200 \times 2.98 \times 4.18}{1000 \times (5.3 \times 10^{-3})}$ $\Delta H = -\ 474$ kJ mol^{-1}	1 1 1 1
	(b)	Assumes that all heat is transferred to the water inside the calorimeter. Assumes that the maximum or minimum temperature is reached before any heat has been lost or gained.	1 1
	(c)	The assumptions and approximations made in calorimetry introduce sources of systematic errors. All heat is transferred to the surroundings The maximum or minimum temperature is reached before any heat has been lost or gained from the atmosphere. The specific heat capacity of any solution is the same as water (4.18 J g^{-1} °C^{-1}) The density of any solution is the same as water (1.0 g mL^{-1})	1 1 (for either) 1 (for either)
	(d) i)	3	1
	ii)	4	1
	iii)	1	1
	iv)	5	1

Summary Test 4: Mixtures and Solutions

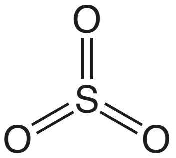
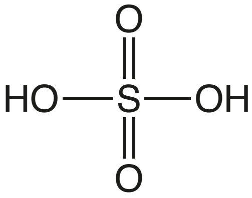
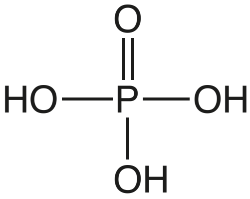
1	(a)	To determine between two unknown substances (one which is known to be cyclohexane, the other ethanol), a simple experiment can be conducted using water. Place 5 mL of each unknown into a large test tube. Into each test tube add 5 mL of water and shake to mix. Based on the principle of miscibility, the sample which is ethanol will form a homogenous solution with water. The sample which is cyclohexane will not form a miscible solution with water, and thus will have an observable layer (meniscus) allowing identification between the two solutions.	1 1
	(b)	The light blue in the diagram represents the solute, which is the substance mixed with the solvent (dark blue). The test tube on the left hand side indicates a miscible relationship between the solute and solvent, the right hand side an immiscible relationship.  	1 2
	(c)	Upon mixing, water and ethanol exhibit secondary interactions between molecules. This formation of these secondary interactions, allows the solute and solvent to readily mix and subsequently deemed miscible. In this specific example, hydrogen bonds will be formed between the water and ethanol molecules.	1 1

Question	Part	Author's response	Marks
	(d)	A homogenous solution is one in which the solute completely and readily dissolves in the solvent.	1
		For this to occur, the attraction among the particles in their original state must be replaced by attractions as strong, or stronger than those between solute and solvent particles.	1
	(e)	Miscibility is dependent on polarity. As a general rule, polar solutes will dissolve in polar solvents forming a miscible solution, as will non-polar solutes in non-polar solvents.	1
		However, when the polarities of solute and solvents are dissimilar, then as a general rule they do not mix, resulting in immiscible solutions. This is due to what occurs at the molecular level between individual solvent and solute molecules.	1
	(f)		2
	(g)	Larger organic molecules have more carbon atoms and hence, a greater number of sites over which dispersion forces will act.	1
		Molecules within the solvent are simply unable to separate the molecules of 1-hexanol given the strength of the attractive forces operating between molecules. If no separation occurs, insolubility results.	1
2	(a)	When potassium chloride is placed in water, the force of attraction between the ions and the molecules of water are greater in magnitude than the ionic bonds within the lattice.	1
		This results in the ions becoming dissociated from the lattice.	1
		Molecules of water separate anions and cations from the lattice and then surround the ions forming a hydration shell. The hydration shell prevents the ions from recrystallising.	1
	(b)	$KCl_{(s)} + aq \rightarrow K^{+}_{(aq)} + Cl^{-}_{(aq)}$	2
	(c)	The molecules of water become bound to the anions and cations at the surface of the ionic lattice.	1
		As an example, the hydrogen atoms of the water molecules (positive pole) are attracted to the anions (negative pole) within the lattice.	1
		As the force of attraction between the ions and the molecules of water are greater in magnitude than the ionic bonds within the lattice, then the ions become dissociated from the lattice.	1
	(d)	A double replacement reaction occurs when two ionic reactants dissociate and bond with the respective anion or cation from the other reactant.	1
		The ions replace each other based on their charges as either a cation or an anion.	1
	(e)	If the reaction involved is a double replacement reaction such as when silver nitrate and potassium chloride, the products of the replacement are first deduced.	1
		In this case, the two products will be silver chloride and potassium nitrate. Once this is known, the solubility table must be consulted to determine if the products are soluble or in fact insoluble precipitates.	1
	(f)	Hard water is defined as water than contains higher than average levels of magnesium and calcium ions.	1
		This can cause the formation of insoluble soap scum when these ions come into contact with soap molecules.	1
	(g)	When the dissociated calcium ions come into contact with the carboxylate anion of soap ions, they are attracted to one another and bond. $Ca^{2+}_{(aq)} + 2RCOO^{-}_{(aq)} \rightarrow Ca(RCOO)_{2(s)}$	1
		The result is a salt that is no longer water-soluble, and it is this precipitate that is commonly called soap scum.	1
		Scum leaves unsightly deposits on clothes, washing machines and in hair. In addition to this, the reaction renders the soap ions unavailable to form grease micelles to assist in the desired cleaning action.	1

Question	Part	Author’s response	Marks
3	(a)	$n_{(MgCO_3)} = c \times v$ $n_{(MgCO_3)} = 0.300 \times 0.03$ $n_{(MgCO_3)} = 9 \times 10^{-3}\text{mol}$	1 1
	(b) (i)	$n_{(MgCO_3)} = \frac{m}{M}$ $n_{(MgCO_3)} = \frac{1.24}{84.31}$ $n_{(MgCO_3)} = 1.47 \times 10^{-2}\text{mol}$ $c_{(MgCO_3)} = \frac{1.47 \times 10^{-2}}{0.2}$ $c_{(MgCO_3)} = 7.35 \times 10^{-2}\text{mol L}^{-1}$	1 1 1 1
	(ii)	$\text{g L}^{-1} = \text{mol L}^{-1} \times M$ $\text{g L}^{-1} = 7.35 \times 10^{-2} \times 84.31$ $= 6.2\ \text{g L}^{-1}$	1+1 1
	(iii)	$\text{ppm} = \text{g L}^{-1} \times 10^3$ $\text{ppm} = 6.2 \times 10^3$ $= 6200\ \text{ppm}$	1 1
	(c)	Representing mass concentration as g L^{-1} is not always the most relevant unit for communicating the mass concentration of a solution. With this in mind other conventions such as percent weight per volume (%w/V), parts per million (ppm) and parts per billion (ppb) can be investigated and units can be interchanged using simple calculations.	1 1
	(d)	A standard solution is a solution of accurately known concentration. It is prepared from the primary standard (a compound which is stable, of high purity, highly soluble in water and of a high molar mass to allow for accurate weighing) that is weighed accurately and made up to a fixed volume.	1 1
	(e)	When reactions are carried out in a laboratory, quantities are measured in grams, not moles. It is for this reasons that most calculations start and finish with mass. This is specifically termed mass-mass stoichiometry. Mass-volume stoichiometry is when either the mass of a substance is given and the volume of another is calculated, or the volume of a substance will be provided, and the mass of another compound will be determined.	1 1
	(f)	$c_iV_i = c_fV_f$ $V_i = \frac{c_f.V_f}{c_i}$ $V_i = \frac{0.4 \times 0.3}{1}$ $V_i = 0.12\ \text{L}\ (120\ \text{mL})$	1 1
4	(a)	$KNO_{3(s)} + aq \rightarrow K^+_{(aq)} + NO^-_{3(aq)}\ \Delta H_{(solution)} = +35\ \text{kJ mol}^{-1}$	4
	(b)	If the products have more energy that the reactants, ΔH will be positive (endothermic) Energy will be absorbed during the reaction. Therefore ΔH can be expressed as: $\Delta H = H_{products} - H_{reactants}$	1 1 1
	(c)	The surroundings refer to water if the reaction takes place in an aqueous environment. Air if it is a combustion reaction or reaction in air, or simply the reaction vessel.	1 1
	(d)	Both lattice energies and hydration energies of ionic compounds are known values therefore can be sourced to determine the overall enthalpy of solution. Essentially the enthalpy of solution can be determined via the following expression: $\Delta H_{solution}$ = lattice energy – hydration energy	1 1

Topic 5: Acids and Bases

Question	Part	Author's response	Marks
1		$HNO_{3(aq)} + H_2O_{(l)} \rightarrow NO^-_{3(aq)} + H_3O^+_{(aq)}$ An acid is a proton donor, nitric acid donates a proton (H^+) to water in this example.	2 1
2		$NaOH_{(s)} + H^+_{(aq)} \rightarrow Na^+_{(aq)} + H_2O_{(l)}$ A base is a proton acceptor, sodium hydroxide accepts a proton from an acid (H^+) in this example.	2 1
3	(a)	CH_3COOH – acid, CH_3COO^- - conjugate base H_2O – base, H_3O^+ - conjugate acid	1 1
	(b)	CO_3^{2-} - base, HCO_3^- - conjugate acid H_2O – acid, OH^- - conjugate base	1 1
	(c)	H_2SO_4 – acid, HSO_4^- - conjugate base NH_3 – base, NH_4^+ - conjugate acid	1 1
	(d)	KOH – base, H_2O – conjugate acid HCl – acid, KCl – conjugate base	1 1
4	(a)	$HCl_{(aq)} + H_2O_{(l)} \rightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$ HCl – acid, Cl^- – conjugate base H_2O – base, H_3O^+ - conjugate acid	2 1 1
	(b)	$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$ HCl – acid, NaCl – conjugate base NaOH – base, H_2O – conjugate acid	2 1 1
	(c)	$CH_3COOH_{(aq)} + NH_{3(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + NH^+_{4(aq)}$ CH_3COOH – acid, CH_3COO^- - conjugate base NH_3 – base, NH_4^+ - conjugate acid	2 1 1
	(d)	$NH^+_{4(aq)} + H_2O_{(l)} \rightarrow NH_{3(aq)} + H_3O^+_{(aq)}$ NH_4^+ - acid, NH_3 – conjugate base H_2O – base, H_3O^+ - conjugate acid	2 1 1
5		$H_2SO_{4(aq)} + 2H_2O_{(l)} \rightarrow SO^{2-}_{4(aq)} + 2H_3O^+_{(aq)}$ Strong acids completely ionise in water, hence the solution only contains sulfate and hydronium ions. $H_2CO_{3(aq)} + 2H_2O(l) \rightleftharpoons CO^{2-}_{3(aq)} + 2H_3O^+_{(aq)}$ Weak acids partially ionise in water hence, the solution will have a greater proportion of carbonic acid molecules to carbonate ions.	2 1 2 1
6		An acid which transfers one proton to a base in a chemical reaction is termed a monoprotic acid. $NH^+_{4(aq)} + H_2O_{(l)} \rightarrow NH_{3(aq)} + H_3O^+_{(aq)}$ The ammonium ion transfers one proton to a molecule of water (base).	1 1 1
7		$H_2S_{(aq)} + H_2O_{(l)} \rightarrow HS^-_{(aq)} + H_3O^+_{(aq)}$ $HS^-_{(aq)} + H_2O_{(l)} \rightarrow S^{2-}_{(aq)} + H_3O^+_{(aq)}$ Hydrogen sulfide is a diprotic acid as it transfers two protons.	2 2 1
8		$H_3BO_{4(aq)} + H_2O_{(l)} \rightarrow H_2BO^-_{4(aq)} + H_3O^+_{(aq)}$ $H_2BO^-_{4(aq)} + H_2O_{(l)} \rightarrow HBO^{2-}_{4(aq)} + H_3O^+_{(aq)}$ $HBO^{2-}_{4(aq)} + H_2O_{(l)} \rightarrow BO^{3-}_{4(aq)} + H_3O^+_{(aq)}$	2 2 2
9		Acids which transfer up to two protons are termed diprotic acids. $H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons HCO^-_{3(aq)} + H_3O^+_{(aq)}$ $HCO^-_{3(aq)} + H_2O_{(l)} \rightleftharpoons CO^{2-}_{3(aq)} + H_3O^+_{(aq)}$	1 2 2

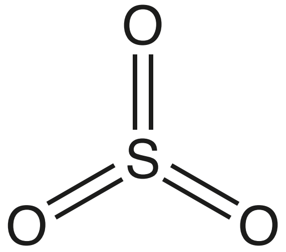
Question	Part	Author's response	Marks
10	(a)	Basic	1
	(b)	Acidic	1
	(c)	Neutral	1
11	(a)	$1s^22s^22p^63s^23p^4$	2
	(b)	$S^{2-}_{(aq)} + 2H_2O_{(l)} \rightarrow H_2S_{(g)} + 2OH^-_{(aq)}$ A basic substance accepts protons from an acid in a chemical reaction.	2 1
	(c)	Blue	1
12	(a)	$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$	2
	(b)	$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$	2
	(c)	$Cl_2O_{(g)} + H_2O_{(l)} \rightarrow 2HOCl_{(aq)}$	2
13	(a)	$SO_{2(g)} + 2NaOH_{(aq)} \rightarrow Na_2SO_{3(aq)} + H_2O_{(l)}$ $SO_{2(g)} + 2OH^-_{(aq)} \rightarrow SO^{2-}_{3(aq)} + H_2O_{(l)}$	2 2
	(b)	$P_4O_{10(s)} + 12NaOH_{(aq)} \rightarrow 4Na_3PO_{4(aq)} + 6H_2O_{(l)}$ $P_4O_{10(s)} + 12OH^-_{(aq)} \rightarrow 4PO^{3-}_{4(aq)} + 6H_2O_{(l)}$	2 2
	(c)	$Cl_2O_{(g)} + 2NaOH_{(aq)} \rightarrow 2NaClO_{(aq)} + H_2O_{(l)}$ $Cl_2O_{(g)} + 2OH^-_{(aq)} \rightarrow 2ClO^-_{(aq)} + H_2O_{(l)}$	2 2
14	(a)		2
	(b)	$O{=}C{=}O$	2
	(c)		2
	(d)		2
15	(a)	$K_2O_{(s)} + H_2O_{(l)} \rightarrow 2K^+_{(aq)} + 2OH^-_{(aq)}$	2
	(b)	$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca^{2+}_{(aq)} + 2OH^-_{(aq)}$	2
	(c)	$Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2Na^+_{(aq)} + 2OH^-_{(aq)}$	2
16	(a)	$Na_2O_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$ $Na_2O_{(s)} + 2H^+_{(aq)} \rightarrow 2Na^+_{(aq)} + H_2O_{(l)}$	2 2
	(b)	$CuO_{(s)} + H_2SO_{4(aq)} \rightarrow CuSO_{4(aq)} + H_2O_{(l)}$ $CuO_{(s)} + 2H^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + H_2O_{(l)}$	2 2
	(c)	$SrO_{(s)} + H_2SO_{4(aq)} \rightarrow SrSO_{4(aq)} + H_2O_{(l)}$ $SrO_{(s)} + 2H^+_{(aq)} \rightarrow Sr^{2+}_{(aq)} + H_2O_{(l)}$	2 2

Question	Part	Author's response	Marks
17	(a)	Elemental oxides that have properties characteristic of both acid and base are amphoteric.	1
		An amphoteric oxide can react as both an acid and a base.	1
	(b)(i)	$ZnO_{(s)} + 2H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_2O_{(l)}$	2
	(ii)	$ZnO_{(s)} + 2OH^{-}_{(aq)} \rightarrow ZnO_{2(aq)}^{2-} + H_2O_{(l)}$	2
18		Increases from left to right.	1
19	(a)	Basic	1
	(b)	Amphoteric	1
	(c)	Acidic	1
20	(a)	$ZnO_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_2O_{(l)}$	2
	(b)	$Al(OH)_{3(s)} + 3HCl_{(aq)} \rightarrow AlCl_{3(aq)} + 3H_2O_{(l)}$	2
	(c)	$Mg(HCO_3)_{2(s)} + 2HNO_{3(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2CO_{2(g)} + 2H_2O_{(l)}$	2
	(d)	$K_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2KCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	2
21	(a)	$KOH_{(aq)} + HNO_{3(aq)} \rightarrow KNO_{3(aq)} + H_2O_{(l)}$ $OH^{-}_{(aq)} + H^{+}_{(aq)} \rightarrow H_2O_{(l)}$	2 2
	(b)	$Na_2CO_{3(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + CO_{2(g)} + H_2O_{(l)}$ $Na_2CO_{3(s)} + 2H^{+}_{(aq)} \rightarrow 2Na^{+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	2 2
	(c)	$KHCO_{3(s)} + HCl_{(aq)} \rightarrow KCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$ $KHCO_{3(s)} + H^{+}_{(aq)} \rightarrow K^{+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	2 2
22		$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$	2
		$v_{(NaOH)} = 50.0\ mL$ $v_{(NaOH)} = 0.0500\ L$ $c_{(NaOH)} = 0.0612\ mol\ L^{-1}$ $n_{(NaOH)} = c_{(NaOH)} \times v_{(NaOH)}$ $n_{(NaOH)} = 0.0612 \times 0.0500$ $n_{(NaOH)} = 0.00306\ mol\ (3.06 \times 10^{-3}\ mol)$	2
		NaOH : HCl 1 : 1 $n_{(NaOH)} = n_{(HCl)}$ $n_{(HCl)} = 0.00306\ mol\ (3.06 \times 10^{-3}\ mol)$	1
		$c_{(HCl)} = 0.103\ mol\ L^{-1}$ $v_{(HCl)} = \frac{n(HCl)}{c(HCl)}$ $v_{(HCl)} = \frac{0.00306}{0.103}$ $v_{(HCl)} = 0.0297\ L\ (3\ s.f.)\ (2.97 \times 10^{-2}\ L)$	2

Question	Part	Author's response	Marks
23		$Na_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	2
		$v_{(HCl)} = 24.55$ mL $v_{(HCl)} = 0.02455$ L $c_{(HCl)} = 0.1034$ mol L^{-1} $n_{(HCl)} = c_{(HCl)} \times v_{(HCl)}$ $n_{(HCl)} = 0.1034 \times 0.02455$ $n_{(HCl)} = 0.002538$ mol (2.538×10^{-3} mol)	2
		Na_2CO_3 : HCl 1 : 2 $n_{(Na_2CO_3)} = \frac{n(HCl)}{2}$ $n_{(Na_2CO_3)} = \frac{0.002538}{2}$ $n_{(Na_2CO_3)} = 0.001269$ mol (1.269×10^{-3} mol)	2
		$v_{(Na_2CO_3)} = 25.0$ mL $v_{(Na_2CO_3)} = 0.0250$ L $c_{(Na_2CO_3)} = \frac{n_{(Na_2CO_3)}}{v_{(Na_2CO_3)}}$ $c_{(Na_2CO_3)} = \frac{0.001269}{0.0250}$ $c_{(Na_2CO_3)} = 0.0508$ mol L^{-1} (3 s.f.) (5.08×10^{-2} mol L^{-1})	2
24		$Na_2O_{(s)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + H_2O_{(l)}$	2
		$v_{(H_2SO_4)} = 50.0$ mL $v_{(H_2SO_4)} = 0.0500$ L $c_{(H_2SO_4)} = 2.013$ mol L^{-1} $n_{(H_2SO_4)} = c_{(H_2SO_4)} \times v_{(H_2SO_4)}$ $n_{(H_2SO_4)} = 2.013 \times 0.0500$ $n_{(H_2SO_4)} = 0.10065$ mol	2
		Na_2O : H_2SO_4 1 : 1 $n_{(Na_2O)} = n_{(H_2SO_4)}$ $n_{(Na_2O)} = 0.10065$ mol	1
		$M_{(Na_2O)} = (M_{Na} \times 2) + (M_O \times 1)$ $M_{(Na_2O)} = (22.99 \times 2) + (16.00)$ $M_{(Na_2O)} = 61.98$ g mol^{-1} $m_{(Na_2O)} = n_{(Na_2O)} \times M_{(Na_2O)}$ $m_{(Na_2O)} = 0.10065 \times 61.98$ $m_{(Na_2O)} = 6.24$ g (3 s.f.)	2

Question	Part	Author's response	Marks
25	(a)	$Na_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	2
	(b)	$v_{(Na_2CO_3)} = 30.0$ mL $v_{(Na_2CO_3)} = 0.0300$ L $c_{(Na_2CO_3)} = 0.0522$ mol L^{-1} $n_{(Na_2CO_3)} = c_{(Na_2CO_3)} \times v_{(Na_2CO_3)}$ $n_{(Na_2CO_3)} = 0.0522 \times 0.0300$ $n_{(Na_2CO_3)} = 0.001566$ mol (1.566×10^{-3} mol)	2
		Na_2CO_3 : NaCl 1 : 2 $n_{(NaCl)} = n_{(Na_2CO_3)} \times 2$ $n_{(NaCl)} = 0.001566 \times 2$ $n_{(NaCl)} = 0.003132$ mol (3.132×10^{-3} mol)	2
		$M_{(NaCl)} = (M_{Na} \times 1) + (M_{Cl} \times 1)$ $M_{(NaCl)} = (22.99) + (35.45)$ $M_{(NaCl)} = 58.44$ g mol^{-1} $m_{(NaCl)} = n_{(NaCl)} \times M_{(NaCl)}$ $m_{(NaCl)} = 0.003132 \times 58.44$ $m_{(NaCl)} = 0.183$ g (3 s.f.)	2
	(c)	Place the solution into an evaporating dish.	1
		Evaporate the liquid off and a crystalline white solid (NaCl) will remain.	1
	(d)	Na_2CO_3 : CO_2 1 : 1 $n_{(CO_2)} = n_{(Na_2CO_3)}$ $n_{(CO_2)} = 0.001566$ mol (1.566×10^{-3} mol)	1
		$M_{(CO_2)} = (M_C \times 1) + (M_O \times 2)$ $M_{(CO_2)} = (12.01) + (16.00 \times 2)$ $M_{(CO_2)} = 44.01$ g mol^{-1} $m_{(CO_2)} = n_{(CO_2)} = \times M_{(CO_2)}$ $m_{(CO_2)} = 0.001566 \times 44.01$ $m_{(CO_2)} = 0.0689$ g (3 s.f.) (6.89×10^{-2} g)	2
26	(a)	Acidic	1
	(b)	Basic	1
	(c)	Neutral	1
27		$[OH^-] = 0.015$ mol L^{-1} $1 \times 10^{-14} = [OH^-].[H_3O^+]$ $[H_3O^+] = \frac{1 \times 10^{-14}}{[OH^-]}$ $[H_3O^+] = \frac{1 \times 10^{-14}}{0.015}$ $[H_3O^+] = 6.7 \times 10^{-13}$ mol L^{-1}	2

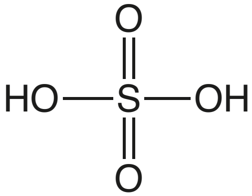
Question	Part	Author's response	Marks
28		$[OH^-] = 0.123 \text{ mol L}^{-1}$ $1 \times 10^{-14} = [OH^-].[H_3O^+]$ $[OH^-] = \frac{1 \times 10^{-14}}{[H_3O^+]}$ $[OH^-] = \frac{1 \times 10^{-14}}{0.123}$ $[OH^-] = 8.13 \times 10^{-14} \text{ mol L}^{-1}$	2
29		$HNO_{3(aq)} + H_2O_{(l)} \longrightarrow NO^-_{3(aq)} + H_3O^+_{(aq)}$ $pH = -\log[H_3O^+]$ $pH = -\log(0.0023)$ $pH = 2.6$	2
30		$HCl_{(aq)} + H_2O_{(l)} \longrightarrow Cl^-_{(aq)} + H_3O^+_{(aq)}$ $[H_3O^+] = 10^{-pH}$ $[H_3O^+] = 10^{-2.1}$ $[H_3O^+] = 0.0079 \text{ mol L}^{-1}$ $(7.9 \times 10^{-3} \text{ mol L}^{-1})$	2
31		$KOH_{(aq)} + aq \longrightarrow K^+_{(aq)} + OH^-_{(aq)}$ $[OH^-] = 10^{-4} \text{ mol L}^{-1}$ $1 \times 10^{-14} = [OH^-].[H_3O^+]$ $[H_3O^+] = \frac{1 \times 10^{-14}}{[OH^-]}$ $[H_3O^+] = \frac{1 \times 10^{-14}}{10^{-4}}$ $[H_3O^+] = 10^{-10} \text{ mol L}^{-1}$ $pH = -\log[H_3O^+]$ $pH = -\log(10^{-10})$ $pH = 10$	2 2
32		$[H_3O^+] = 10^{-pH}$ $[H_3O^+] = 10^{-8.3} \text{ mol L}^{-1}$ $[OH^-] = \frac{1 \times 10^{-14}}{[H_3O^+]}$ $[OH^-] = \frac{1 \times 10^{-14}}{10^{-8.3}}$ $[OH^-] = 10^{-5.7} \text{ mol L}^{-1}$	2 2

Question	Part	Author's response	Marks
33	(a) i)	Sulfur is found as a contaminant in oil and fuel. OR Sulfur is an impurity in metallic ores.	1
		When these are burnt in oxygen, sulfur and oxygen react to form sulfur dioxide	1
	ii)		2
		Trigonal planar	1
	(b) i)	Acid rain is defined as rainwater with a pH less than 5.6. Acid rain occurs when the hydronium ion concentration in water droplets is increased.	2
	ii) (1)	$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$	2
	(2)	$2NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)} + HNO_{2(aq)}$	2
33	(c)	Marks allocated for a **total** of 6 marks	
		The lower pH of rainwater can cause plant leaves to lose their waxy coating.	1
		This leads to the leaves to become discoloured and susceptible to bacteria attacks.	1
		Clay particles in soil are negatively charged and attract metal cations to their surface.	1
		The hydrogen ions (H^+) from the rainwater are able to exchange the metal cations (e.g. K^+) from the surface of the clay.	1
		$K^+_{(clay)} + H^+_{(aq)} \rightleftharpoons K^+_{(aq)} + H^+_{(clay)}$	1
		This results in nutrients being leached away so that plant root systems can no longer use them.	1
		Aluminium ions are leached from soil and flow into waterways.	1
		Once in waterways they can adhere to fish gills reducing their ability to extract oxygen and sometimes resulting in death.	1
		Acid rain reduces the pH levels of water bodies and as a result causes detrimental effects to the aquatic life.	1
		When the pH of the water reaches 5 fish eggs cannot hatch and lower pH levels cause some adult fish to die.	1
		Calcium carbonate is an insoluble compound, however when it reacts with strong acids (H^+) in acid rain more soluble calcium salts (Ca^{2+}) are formed.	1
		$CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	1
		The soluble calcium salts are then washed away with rainwater, leading to the eroding of limestone in the environment and of man-made structures throughout the world.	1
		Humans can experience trouble with their respiratory system.	1
		Fine sulfate and nitrate particles produced from interactions in the atmosphere can be transported over long distances and then are inhaled by humans into their lungs.	1

Summary Test 5: Acids and Bases

Question	Part	Author's response	Marks
1	(a)	$H_2CO_{3(aq)} \rightleftharpoons CO_{2(aq)} + H_2O_{(aq)}$ Carbon dioxide is released.	2 1
	(b)	A weak acid partially ionises in water. When one mole of carbonic acid reacts with water it produces less than one mole of hydronium ions.	1 1
	(c)	Hydrochloric acid is a strong acid and will completely ionise in solution. There will be a higher concentration of hydronium ions in the hydrochloric acid solution therefore, the solution will have a lower pH than the carbonic acid solution.	1 1
	(d)	Hydrochloric acid would have a greater electrical conductivity. Higher concentration of charged ions in solution.	1 1
	(e)	Carbon dioxide is produced. Carbon dioxide would extinguish the flame.	1 1
	(f)	$H_2CO_{3(aq)} + H_2O_{(l)} \rightleftharpoons HCO^-_{3(aq)} + H_3O^+_{(aq)}$ $HCO^-_{3(aq)} + H_2O_{(l)} \rightleftharpoons CO^{2-}_{3(aq)} + H_3O^+_{(aq)}$	2 2
	(g)	Red	1
	(h)	$m_{(CO_2)} = 2.401$ g $M_{(CO_2)} = (M_C \times 1) + (M_O \times 2)$ $M_{(CO_2)} = (12.01) + (16.00 \times 2)$ $M_{(CO_2)} = 44.01$ g mol^{-1} $n_{(CO_2)} = \frac{m_{(CO_2)}}{M_{(CO_2)}}$ $n_{(CO_2)} = \frac{2.401}{44.01}$ $n_{(CO_2)} = 0.05456$ mol (5.456×10^{-2} mol) $H_2CO_3 : CO_2$ 1 : 1 $n_{(H_2CO_3)} = n_{(CO_2)}$ $n_{(H_2CO_3)} = 0.05456$ mol (5.456×10^{-2} mol) $M_{(H_2CO_3)} = (M_H \times 2) + (M_C \times 1) + (M_O \times 3)$ $M_{(H_2CO_3)} = (1.008 \times 2) + (12.01 \times 1) + (16.00 \times 3)$ $M_{(H_2CO_3)} = 62.026$ g mol^{-1} $m_{(H_2CO_3)} = n_{(H_2CO_3)} \times M_{(H_2CO_3)}$ $m_{(H_2CO_3)} = 0.05456 \times 62.026$ $m_{(H_2CO_3)} = 3.384$ g (3 s.f.)	2 1 2

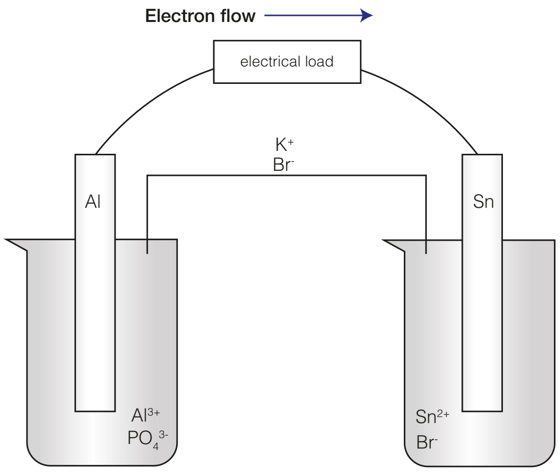
Question	Part	Author's response	Marks
2	(a)	$\rho_{(NH_3)} = 20\text{ g L}^{-1}$ $M_{(NH_3)} = (M_N \times 1) + (M_H \times 3)$ $M_{(NH_3)} = (14.01 \times 1) + (1.008 \times 3)$ $M_{(NH_3)} = 17.034\text{ g mol}^{-1}$ $c_{(NH_3)} = \frac{\rho_{(NH_3)}}{M_{(NH_3)}}$ $c_{(NH_3)} = \frac{20}{17.034}$ $c_{(NH_3)} = 1.174\text{ mol L}^{-1}$	2
	(b)	$c_1V_1 = c_2V_2$ $1.174 \times 0.025 = c_2 \times 0.25$ $c_2 = \frac{1.174 \times 0.025}{0.25}$ $c_2 = 0.1174\text{ mol L}^{-1}$	2
	(c)	The flask was inverted to ensure the solution was homogenous.	1
	(d)	$NH_{3(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)}$	2
	(e)	Concordant titres are within 0.1 mL of each other.	1
	(f)	$\frac{23.65 + 23.60 + 23.65}{0.25} = 23.63\text{ mL}$	1
	(g)	$V_{(HCl)} = 0.02363\text{ L}$ $c_{(HCl)} = 0.0959\text{ mol L}^{-1}$ $n_{(HCl)} = c_{(HCl)} \times V_{(HCl)}$ $n_{(HCl)} = 0.0959 \times 0.02363$ $n_{(HCl)} = 0.002266\text{ mol } (2.266 \times 10^{-3}\text{ mol})$ NH_3 : HCl 1 : 1 $n_{(NH_3)} = n_{(HCl)}$ $n_{(NH_3)} = 0.002266\text{ mol } (2.266 \times 10^{-3}\text{ mol})$ $V_{(NH_3)} = 0.0250\text{ L}$ $c_{(NH_3)} = \frac{n_{(NH_3)}}{V_{(NH_3)}}$ $c_{(NH_3)} = \frac{0.002266}{0.0250}$ $c_{(NH_3)} = 0.0906\text{ mol L}^{-1}$ (3 s.f.) $(9.06 \times 10^{-2}\text{ mol L}^{-1})$	2 1 2
	(h)	$c_1V_1 = c_2V_2$ $c_1 \times 0.025 = 0.0906 \times 0.25$ $c_1 = \frac{0.0906 \times 0.25}{0.25}$ $c_2 = 0.906\text{ mol L}^{-1}$	2
	(i)	No it was lower. The claimed molar concentration was 1.174 mol L^{-1}.	1
3	(a)	Acid rain is defined as rainwater with a pH less than 5.6. Acid rain occurs when the hydronium ion concentration in water droplets is increased.	1 1
	(b)	$CaCO_{3(s)} + 2H^+_{(aq)} \rightarrow Ca^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$	2
	(c)	Increases the likelihood of hard water forming.	1
	(d)	(one of the following) Soil leaching, destroys fish eggs, fish gills or erosion of marble/limestone structures.	1

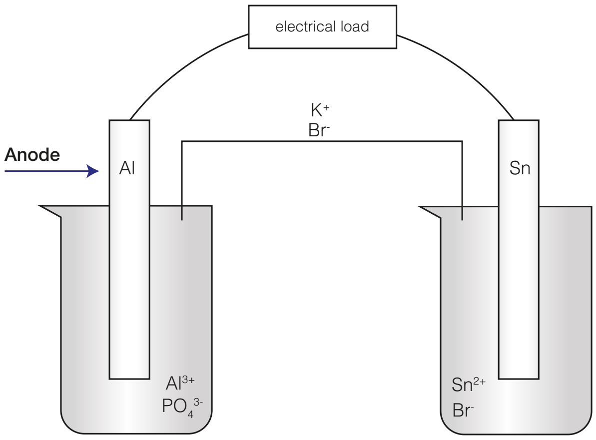
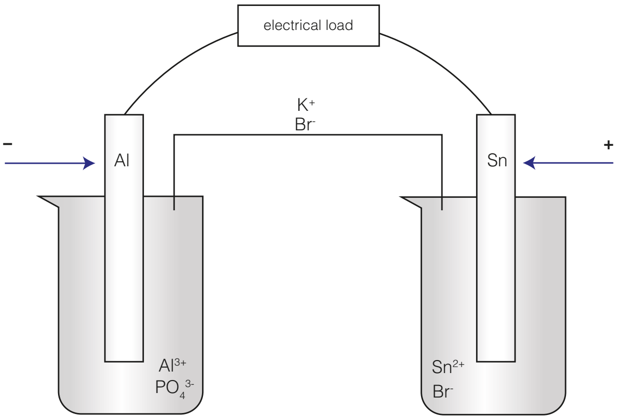
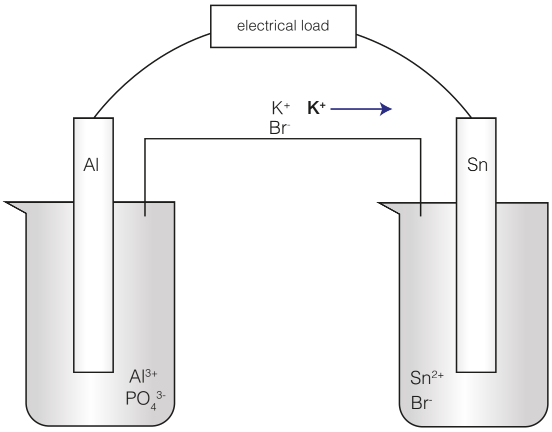
Question	Part	Author's response	Marks
4	(a)	Volcanoes	1
	(b)	$2ZnS_{(s)} + 3O_{2(g)} \rightarrow 2ZnO_{(s)} + 2SO_{2(g)}$	2
	(c)	Note: Equations do not need to be balanced for full marks.	
		$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$	1
		$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$	1
		$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$	1
		The oxides of sulfur dissolve in rainwater forming sulfurous and sulfuric acid.	1
		$H_2SO_{4(aq)} + 2H_2O_{(l)} \rightarrow 2H_3O^+_{(aq)} + SO^{2-}_{4(aq)}$	1
		Sulfuric acid completely ionises in rainwater to increase the concentration of hydronium ions and lower the pH of rainwater.	1
	(d)	 $HO-\overset{\overset{\displaystyle O}{\Vert}}{\underset{\underset{\displaystyle O}{\Vert}}{S}}-OH$	2
	(e)	(one of the following topics explained in full for two marks)	
		Scrubbers in chimneys or catalytic converters in cars.	1
		Scrubbers are used in industry smoke stacks and they neutralise the harmful gases before they are released into the atmosphere.	1
		Catalytic converters are used in cars and they convert harmful gases (e.g. CO and NO) into less harmful gases (e.g. CO_2, N_2 and H_2O).	

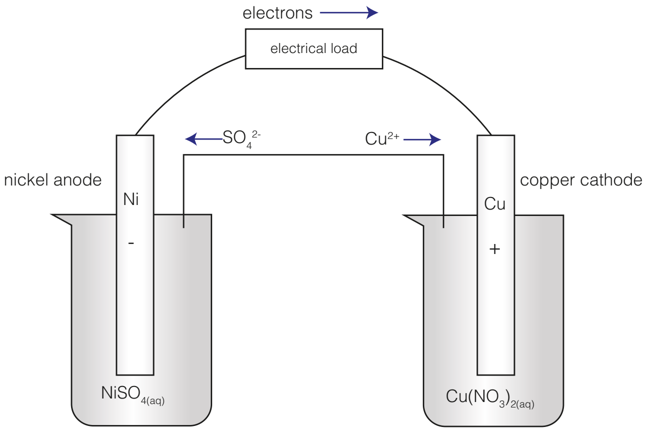
Topic 6: Redox Reactions

Question	Part	Author's response	Marks
1	(a)	+1	1
	(b)	+4	1
	(c)	+7	1
	(d)	+3	1
	(e)	+5	1
	(f)	–1	1
2	(a)	–3	1
	(b)	+3	1
	(c)	+4	1
	(d)	+5	1
	(e)	–2	1
	(f)	+1	1
3	(a)	Li: +1 Cl: –1	2
	(b)	O: 0	1
	(c)	N: –3 H: +1	2
	(d)	Na: +1 H: –1	2
	(e)	C: –2 H: +1 O: –1	3
	(f)	O: +2 F: –1	2
	(g)	H: +1 C: +4 O: –2	3
	(h)	H: +1 S: +6 O: –2	3
	(i)	Al: +3 O: –2 H: +1	3
	(j)	Mn: +4 O: –2	2
4	(a)	Oxidised species: sodium	1
	(b)	Reduced species: hydrogen	1
5	(a)	Oxidising agent: chlorine	1
	(b)	Reducing agent: aluminium	1
6	(a)	Lost electrons: iron	1
	(b)	Gained electrons: oxygen	1
7	(a)	"oxidation"	1
	(b)	"Redox" or "Reduction-oxidation"	1
	(c)	"reduced"; "oxidising"	2
	(d)	"reduced"	1

Question	Part	Author’s response	Marks
8	(a)	The oxidation number of oxygen has changed from 0 to –2. The oxidation number has decreased by 2.	1
	(b)	Oxygen has been reduced.	1
	(c)	The oxidation number of phosphorus has changed from 0 to +5. The oxidation number has increased by 5.	1
	(d)	$P_4 \rightarrow 4P^{5+} + 20e^-$	1
	(e)	**Any one of:** Electrons are present on the product side. Phosphorus has acquired a positive charge.	1
9	(a)	In CO_2: C: +4 O: –2 In H_2O: H: +1 O: –2 In $C_6H_{12}O_6$: C: 0 H: +1 O: –2 In O_2: O: 0	1 1 1 1
	(b)	Oxygen has been oxidised. (Its oxidation number increased from –2 to 0).	1
	(c)	$6CO_2 + 24H^+ + 24e^- \rightarrow C_6H_{12}O_6 + 6H_2O$	2
10		Reduction is a loss of electrons; **or** the species that is reduced transfers one or more electrons to another species.	1
11	(a)	The oxidation number of bromine has changed from –1 to 0. The oxidation number has increased by 1.	1
	(b)	Oxidised	1
	(c)	$2Br^- \rightarrow Br_2 + 2e^-$	2
	(d)	The oxidation number of chlorine has changed from 0 to –1. The oxidation number has decreased by 1.	1
12	(a)	$Al \rightarrow Al^{3+} + 3e^-$ $I_2 + 2e^- \rightarrow 2I^-$	2 2
	(b)	$(Al \rightarrow Al^{3+} + 3e^-) \times 2 =$ $2Al \rightarrow 2Al^{3+} + 6e^-$ $(I_2 + 2e^- \rightarrow 2I^-) \times 3 =$ $3I_2 + 6e^- \rightarrow 6I^-$ $2Al + 3I_2 + 6e^- \rightarrow 2Al^{3+} + 6e^- + 6I^-$ $2Al + 3I_2 \rightarrow 2Al^{3+} + 6I^-$ **or** $2Al + 3I_2 \rightarrow 2AlI_3$	1 1
	(c)	Aluminium	1
13	(a)	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$	1
	(b)	$Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$	1
	(c)	$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$	1
	(d)	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1
	(e)	$CH_3OH + H_2O \rightarrow HCOOH + 4H^+ + 4e^-$	1
	(f)	$(Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O) \times 2 =$ $2Cr_2O_7^{2-} + 28H^+ + 12e^- \rightarrow 4Cr^{3+} + 14H_2O$ $(CH_3OH + H_2O \rightarrow HCOOH + 4H^+ + 4e^-) \times 3 =$ $3CH_3OH + 3H_2O \rightarrow 3HCOOH + 12H^+ + 12e^-$	1 1
	(g)	$2Cr_2O_7^{2-} + 16H^+ + 3CH_3OH \rightarrow 4Cr^{3+} + 11H_2O + 3HCOOH$	2
14	(a)	$CrO_3 + 6H^+ + 3e^- \rightarrow Cr^{3+} + 3H_2O$	2
	(b)	$CH_3CH_2CHO + H_2O \rightarrow CH_3CH_2COOH + 2H^+ + 2e^-$	2
	(c)	$2CrO_3 + 3CH_3CH_2CHO + 6H^+ \rightarrow 2Cr^{3+} + 3CH_3CH_2COOH + 3H_2O$	2

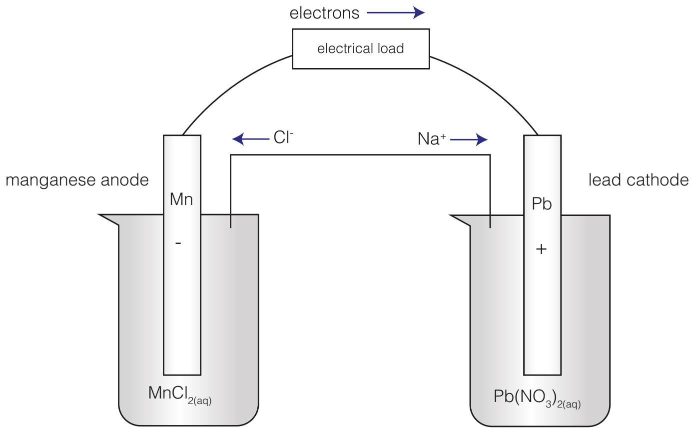
Question	Part	Author's response	Marks
15	(a)	$Cu(NO_3)_2$	1
	(b)	Copper	1
	(c)	$Cu \rightarrow Cu^{2+} + 2e^-$	2
	(d)	$Ag^+ + e^- \rightarrow Ag$	2
	(e)	$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$	1
	(f)	$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$	1
	(g)	The gold wire and solution would remain unchanged.	1
	(h)	Because gold is less reactive than silver, no reaction is expected to occur.	1
16	(a)	A reaction would occur.	1
	(b)	A reaction would not occur.	1
	(c)	A reaction would occur.	1
17	(a)	$Zn + Sn^{2+} \rightarrow Zn^{2+} + Sn$ $Zn + SnCl_2 \rightarrow ZnCl_2 + Sn$	1 2
	(b)	$Sr + Cu^{2+} \rightarrow Sr^{2+} + Cu$ $Sr + CuSO_4 \rightarrow SrSO_4 + Cu$	1 2
	(c)	$2Al + 3Ni^{2+} \rightarrow 2Al^{3+} + 3Ni$ $2Al + 3Ni(NO_3)_2 \rightarrow 2Al(NO_3)_3 + 3Ni$	1 2
18	(a)	$Mn \rightarrow Mn^{2+} + 2e^-$	2
	(b)	$2H^+ + 2e^- \rightarrow H_2$	2
19	(a)	$Co \rightarrow Co^{2+} + 2e^-$	2
	(b)	$2H^+ + 2e^- \rightarrow H_2$	2
	(c)	$Co + 2H^+ \rightarrow Co^{2+} + H_2$	2
	(d)	$3Co + 2H_3PO_4 \rightarrow Co_3(PO_4)_2 + 3H_2$	2
20	(a)	Aluminium	1
	(b)		1
	(c)	Oxidised	1

Question	Part	Author's response	Marks
20	(d)		1
	(e)		1
	(f)	$Al \rightarrow Al^{3+} + 3e^-$	2
	(g)	$Sn^{2+} + 2e^- \rightarrow Sn$	2
	(h)	$(Al \rightarrow Al^{3+} + 3e^-) \times 2 = \; 2Al \rightarrow 2Al^{3+} + 6e^-$ $(Sn^{2+} + 2e^- \rightarrow Sn) \times 3 = 3Sn^{2+} + 6e^- \rightarrow 3Sn$ $2Al + 3Sn^{2+} \rightarrow 2Al^{3+} + 3Sn$	3
	(i)	**Any one of:** Completing the circuit Equilibrating charge Allowing ions to flow between the two solutions	1
	(j)		1

Question	Part	Author's response	Marks
21	(a)	 **One mark each for:** The diagram shows two electrodes, each in an electrolyte solution, with a connecting wire and salt bridge. 'Anode' and 'cathode' labels are present. Nickel is labelled as the anode, and copper is labelled as the cathode. The anode is negative, and the cathode is positive. Electrons flow from nickel to copper through the wire. Anions (SO_4^{2-}) flow towards the Ni/Ni^{2+} half-cell, and cations (Cu^{2+}) flow to the Cu/Cu^{2+} half-cell.	6
	(b)	Nickel	1
	(c)	$Ni \rightarrow Ni^{2+} + 2e^-$	2
	(d)	$Cu^{2+} + 2e^- \rightarrow Cu$	2
	(e)	$Ni + Cu^{2+} \rightarrow Ni^{2+} + Cu$	2
22	(a)	Lead's oxidation number has increased from 0 to +2.	1
	(b)	Oxidised	1
	(c)	Anode	1
	(d)	$PbO_2 + SO_4^{2-} \rightarrow PbSO_4$ $PbO_2 + SO_4^{2-} \rightarrow PbSO_4 \mathbf{+ 2H_2O}$ $PbO_2 + SO_4^{2-} \mathbf{+ 4H^+} \rightarrow PbSO_4 + 2H_2O$ $PbO_2 + SO_4^{2-} + 4H^+ \mathbf{+ 2e^-} \rightarrow PbSO_4 + 2H_2O$	1 1 1 1
	(e)	The sulfate ions are **not** spectator ions.	1
	(f)	The sulfate ions are **not** spectator ions.	1

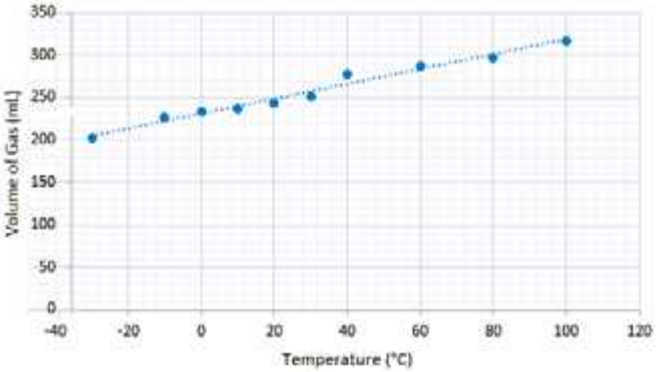
Summary Test 6: Redox Reactions

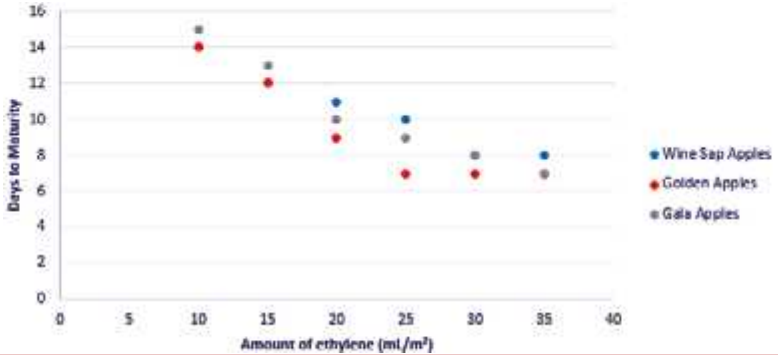
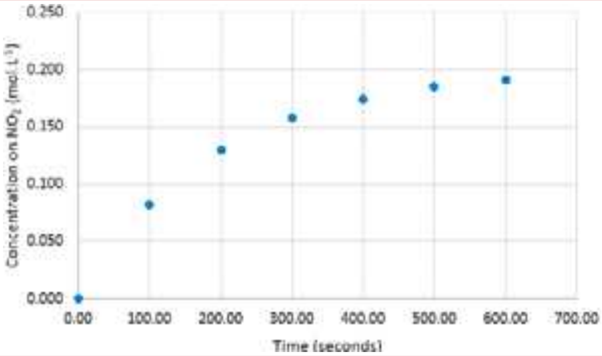
Question	Part	Author’s response	Marks
1	(a)	In K_2CO_3: K: +1 C: +4 O: –2 In HNO_3: H: +1 N: +5 O: –2 In KNO_3: K: +1 N: +5 O: –2 In H_2O: H: +1 O: –2 In CO_2: C: +4 O: –2	1 1 1 1 1
	(b)	Oxidation is an increase in oxidation number.	1
	(c)	The reaction is not a redox reaction. No atom has changed oxidation number; no species has been oxidised or reduced.	1 1
2	(a)	0	1
	(b)	–1	1
	(c)	$F_2 + 2e^- \rightarrow 2F^-$	2
	(d)	1 electron	1
	(e)	2 electrons	1
	(f)	$Ca \rightarrow Ca^{2+} + 2e^-$	2
3	(a)	$Zn \rightarrow Zn^{2+} + 2e^-$	2
	(b)	$I_2 + 2e^- \rightarrow 2I^-$	2
	(c)	Iodine is the oxidising agent	1
4		$Cu \rightarrow Cu^{2+} + 2e^-$ $HNO_3 + 3H^+ + 3e^- \rightarrow NO + 2H_2O$ $3Cu + 6H^+ + 2HNO_3 \rightarrow 3Cu^{2+} + 2NO + 4H_2O$	2 2 2
5	(a)	$Mg \rightarrow Mg^{2+} + 2e^-$	2
	(b)	$Cu^{2+} + 2e^- \rightarrow Cu$	2
	(c)	$Mg + CuSO_4 \rightarrow MgSO_4 + Cu$	2
	(d)	Copper(II) cations (Cu^{2+}) are the reducing agent.	1
	(e)	To displace something means to move it from its current position. In displacement reactions, more reactive metals *displace* less reactive metals in an ionic compound.	1 1
6		Gold is highly unreactive, so pure gold in nature does not usually react with other substances. Sodium is highly reactive. It will react with moisture in the air to produce sodium hydroxide, which is extremely corrosive, and hydrogen, which is flammable. Storing sodium metal under mineral oil prevents it from encountering water or other potential oxidising agemts.	1 1

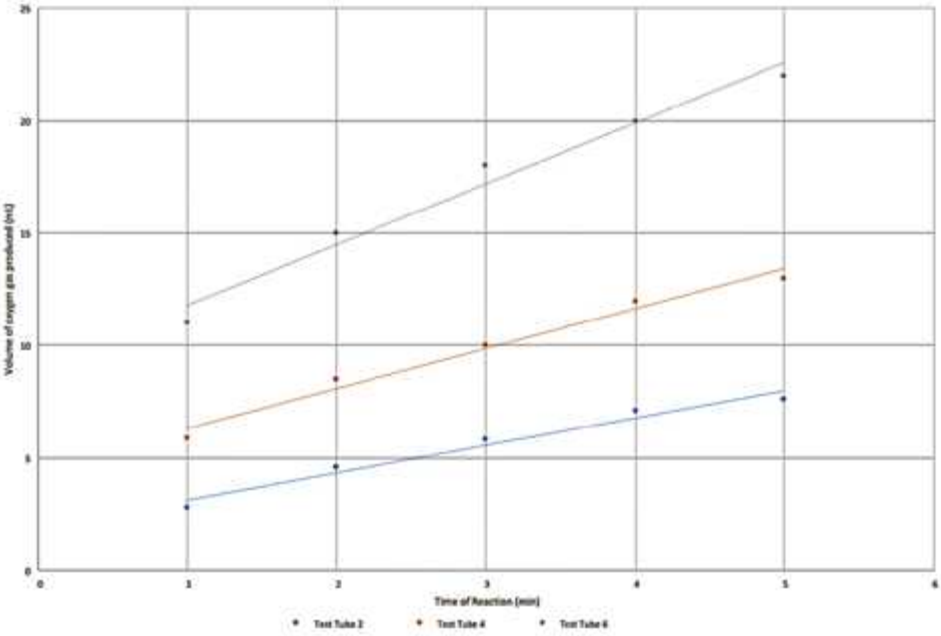
Question	Part	Author’s response	Marks
7		 **One mark each for:** The diagram shows two electrodes, each in an electrolyte solution, with a connecting wire and salt bridge. ‘Anode’ and ‘cathode’ labels are present. Manganese is labelled as the anode, and lead is labelled as the cathode. The anode is negative, and the cathode is positive. Electrons flow from manganese to lead through the wire. Anions (Cl^-) flow towards the Mn/Mn^{2+} half-cell, and cations (Na^+) flow to the Pb/Pb^{2+} half-cell.	
8	(a)	In CoO_2: Co: +4 O: –2 In Li^+: Li: +1 In $LiCoO_2$: Li: +1 Co: +3 O: –2	1 1 1
	(b)	Cobalt	1
	(c)	Cathode	1

Topic 7: Science Inquiry Skills

Question	Part	Author's response	Marks
1	(a)	Clothing failed to look clean after being washed in laundry detergent.	1
	(b)	Shane asked a question based around his problem. His friend Brad provided some background research, which led to the implementation of a hypothesis. This was then tested in an experiment, resulting observations from which a conclusion could be drawn.	1 1 1
	(c)	Prime detergent did not exhibit superior cleaning action to Shane's regular laundry detergent.	1
	(d)	No. The data obtained is a subjective observation, therefore a sound hypothesis cannot be deducted.	1 1
2		Unsound Sound Sound Unsound	1 1 1 1
3	(a)	Type of milk used (with varying fat content)	1
	(b)	Independent	1
	(c)	To investigate the effect of fat content in milk on the functionality of the casein glue that can be produced from its acidification.	1
	(d)	The casein glue produced from milk of higher fat content will have decreased adhesive ability.	1 + 1
	(e)	Use a sample of each of the glues produced to adhere two popsicle sticks (or similar) together. Measure the amount of weight that the adhered popsicle sticks can withstand before breaking.	1 1
	(f)	Volume of acid (vinegar) added Volume of milk added	1 1
	(g)	Polymerisation Acid and Bases	1 1
4	(a)	Temperature	1
	(b)	Mass of glucose that crystallised	1
	(c)	If the temperature of the glucose solution is lowered, the mass of crystallised sugar will increase.	1+1

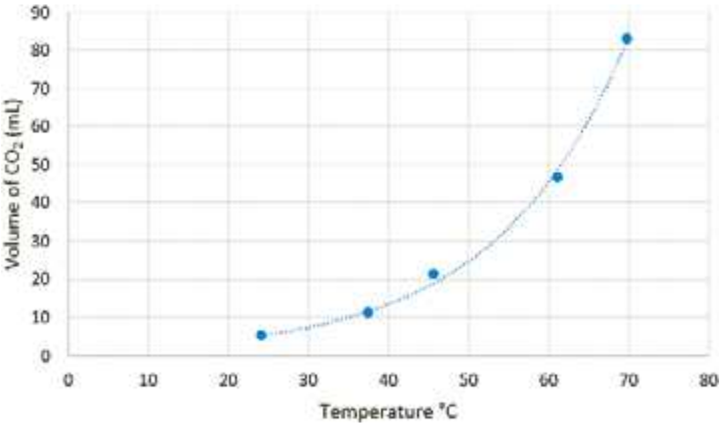
<table>
<tr><th>Question</th><th>Part</th><th>Author’s response</th><th>Marks</th></tr>
<tr><td rowspan="8">5</td><td>(a)</td><td>If the mass of powdered zinc that is added is increased, upon reaction with the hydrochloric acid, the volume of hydrogen gas evolved will also increase.</td><td>1+1</td></tr>
<tr><td>(b)</td><td>Mass of powdered zinc (g)</td><td>1</td></tr>
<tr><td>(c)</td><td>Volume of hydrogen gas evolved (g)</td><td>1</td></tr>
<tr><td>(d)</td><td>Concentration of hydrochloric acid
Method of collecting the gas
Time of reaction</td><td>1
1
1</td></tr>
<tr><td>(e)</td><td>Corrosive nature of acid being used.
To minimise risk wear gloves and keep away from work area when not in use.</td><td>1
1</td></tr>
<tr><td>(f)</td><td>Additional trials to ensure repetition of each length.
This allows an average to be calculated, improving the precision of the experimental data, thus reducing the effect of any random error.</td><td>1
1</td></tr>
<tr><td>(g)</td><td>It is incorrect.
Averaging data improves the precision, as it minimises the effect of any random errors that may be present. Calculating an average does not improve accuracy.</td><td>1
1</td></tr>
<tr><td>(h)</td><td>Constructive and productive collaboration will involve both members of the team being active throughout the practicals and being responsible for particular tasks.
Each member will engage in communication relevant to the practical ensuring they are both aware of what is required for the next step.</td><td>1
1</td></tr>
<tr><td rowspan="2">6</td><td>(a)</td><td>
<table>
<tr><th>Time (min)</th><th>Temperature (°C)</th></tr>
<tr><td>0</td><td>22</td></tr>
<tr><td>1</td><td>24</td></tr>
<tr><td>2</td><td>30</td></tr>
<tr><td>3</td><td>38</td></tr>
<tr><td>4</td><td>46</td></tr>
<tr><td>5</td><td>54</td></tr>
<tr><td>6</td><td>60</td></tr>
<tr><td>7</td><td>64</td></tr>
<tr><td>8</td><td>70</td></tr>
<tr><td>9</td><td>76</td></tr>
<tr><td>10</td><td>82</td></tr>
</table>
</td><td>4</td></tr>
<tr><td>(b)</td><td>Derived.
Degrees Celsius are calculated from the fundamental unit Kelvin (°C = K – 273.15).</td><td>1
1</td></tr>
<tr><td>7</td><td>(a)</td><td>
</td><td>5</td></tr>
</table>

Question	Part	Author's response	Marks
8	(a)	 	5
9	(a)	7.42×10^{-2}	1
	(b)	1.8×10^{-3}	1
	(c)	4×10^{-4}	1
	(d)	-5.8×10^{-1}	1
	(e)	5.064×10^{-1}	1
	(f)	7.006×10^{3}	1
	(g)	-2.1552×10^{4}	1
	(h)	2.15×10^{2}	1
10	(a)	 	5
	(b)	0.162 mol L^{-1}	1
	(c)	0.200 mol L^{-1}	1
	(d)	The value for (b) was obtained as a new data point within the original data set therefore is an example of interpolation. Extrapolation involves predicting values beyond the data set, this was required to be done to obtain the value for (c).	1 1

Question	Part	Author's response	Marks
11	(a)	If the amount of MnO_2 added is increased, the volume of O_2 produced will increase in a proportional manner.	1 + 1
	(b) i)	2.8 mL	1
	ii)	12 mL	1
	iii)	Sum the volume of all test tubes after 5 minutes = 157.7 mL. 0.16 L of O_2 present after 5 minutes.	2
	(c)		5
	(d) i)	9 mL	1
	ii)	11 mL	1
	(e)	Linear	1
	(f)	Increases to the independent variable causes the dependant variable to increase regularly.	1
	(g)	A greater mass of added MnO_2 will result in a greater volume of O_2 being produced when it is reacted with H_2O_2.	1
10	(a)	A	1
	(b)	Thermometer used.	1
11	(a)	Refers to level of reproducibility of measurements. The precision of a practical is dependent on the extent to which random errors can be minimised.	1 1
	(b)	Brad has the greatest scatter (2.9). Shane and Natalie have a lower degree of scatter (0.5).	1 1
	(c)	The average of Shane's results are closest to the true value (22.70). Therefore Shane is more accurate than Brad.	1 1
	(d)	The results of Shane and Natalie are measured to the same resolution (4 significant figures). Brad's results have a lower resolution, being recorded to 3 significant figures.	1 1

Question	Part	Author's response	Marks
12	(a)	Balance 1	1
	(b)	Balance 1	1
	(c)	Balance 2	1
	(d)	Balance 1 has the greatest resolution (5 significant figures). Balance 2 has lower measuring resolution (4 significant figures).	1 1
13		Systematic. The results obtained indicate a calibration error influencing the accuracy of the data.	1 1
14	(a)	5, 3, 4 and 2.	2
	(b)	Accuracy of data refers to the closeness of that data to the known value. In this case, how close the collected data relates to the minimum gross weight of the coin (3.108g). The mint scale produces results closest to this, thus is the most accurate.	1 1
	(c)	Precision refers to the level of agreement in the measurements/data. The Mint scale has the lowest range of data values (0.001) and therefore is the most precise.	1 1

Summary Test 7: Science Inquiry Skills

Question	Part	Author's response	Marks
1	(a)	An increase in temperature will cause a greater volume of carbon dioxide to be produced.	1+1
	(b)	Yes. An increase of 10°C provides almost a doubling of the volume of carbon dioxide.	1 1
	(c)	Volume of hydrochloric acid Concentration of hydrochloric acid Mass of marble chips Volume / size of conical flask	2 (any two)
	(d)	Volume of carbon dioxide was not measured at eye level consistently. This decreases the precision of the data obtained in each of the flasks.	1 1
	(e)	Ensure that the same experimenter is responsible for taking the measurement at eye level, and implementing the same procedure consistently.	1
	(f)	The thermometer used was not calibrated. This resulted in temperature recordings that were inaccurate.	1 1
	(g)	Replace the thermometer with one that is calibrated.	1
	(h)		5
	(i) (i)	7mL	1
	(ii)	26mL	1

Question	Part	Author's response	Marks
2	(a)	To investigate the effect of insulation of the $\Delta H_{neutralisation}$ of nitric acid and sodium hydroxide.	1+1
	(b)	If the insulation of the reaction vessel is improved, then the temperature rise recorded when nitric acid and sodium hydroxide react together will increase.	1+1
	(c)	Temperature change of solution	1
	(d)	Method of insulation of the reaction vessel	1
	(e)	Volume of nitric acid and sodium hydroxide Concentration of nitric acid and sodium hydroxide	1 1
	(f)	0.150 mol L^{-1} sodium hydroxide 0.150 mol L^{-1} nitric acid 200 mL beaker Stirring rod Polystyrene cup Insulated flask 100°C thermometer 2 × 50 mL measuring cylinder	1 (quantities) 1 (materials/apparatus)
	(g)	Corrosive nature of the reactants being used. To minimise their risk ensure dropper bottles are used to deliver volumes of reactant into reaction vessel.	1 1
	(h)	Contamination / purity of sodium hydroxide or nitric acid. This will affect the accuracy of the results produced.	1 1
	(i)	Experiment 3. As the temperature change recorded is the closest to the known value.	1 1
3	(a)	Alcohol (2 significant figures).	1
	(b)	To investigate the effect of increasing the mass of sulfuric acid on the mass of ester produced.	1+1
	(c)	If the mass of sulfuric acid increases, the mass of ester produced will increase.	1+1
	(d)	Mass of sulfuric acid (g).	1
	(e)	Mass of ester produced (g).	1
	(f)	Miscalibration of the mass balance used. This can result in inaccurate measurements of reactants being recorded, affecting the overall accuracy of the practical results.	1 1
	(g)	Each student to complete two trials of the practical to allow an average to be calculated. Two mass balances used to measure reactants to improve accuracy of measurements taken.	1 1
	(h)	The mass of ester produced increases upon the mass of sulfuric acid being increased up to a mass of 3.50g. A mass of sulfuric acid greater than 3.50g resulted in a subsequent decrease in the mass of ester produced.	1 1

Question	Part	Author's response	Marks
4	(a)	A volumetric pipette is calibrated to deliver a particular volume of solution when used correctly.	1
		Thus it is more accurate than a measuring cylinder.	1
	(b)	The burette must first be rinsed with distilled water.	1
		This ensures the burette is free from any contaminants from previous experiments.	1
		Following this, the burette should be rinsed with the solution it is to contain for the experiment.	1
		This ensures that the solution is not diluted from the distilled water rinse.	1
	(c)	This introduces a random error, as the aliquot measured with the pipette which was rinsed only with distilled water would be diluted.	1
		The result of this would be a decreased titre as less moles of permanganate would be required to react with the diluted peroxide.	1

Appendicies

Appendix 1: Molar Masses of the Elements

Name	Symbol	At. No.	At. Mass.
Actinium	Ac	89	227*
Aluminum	Al	13	26.98
Americium	Am	95	243*
Antimony	Sb	51	121.8
Argon	Ar	18	39.95
Arsenic	As	33	74.92
Astatine	At	85	210*
Barium	Ba	56	137.3
Berkelium	Bk	97	245*
Beryllium	Be	4	9.012
Bismuth	Bi	83	208.9
Bohrium**	Bh	107	272
Boron	B	5	10.81
Bromine	Br	35	79.90
Cadmium	Cd	48	112.4
Calcium	Ca	20	40.08
Californium	Cf	98	248*
Carbon	C	6	12.01
Cerium	Ce	58	140.1
Cesium	Cs	55	132.9
Chlorine	C1	17	35.45
Chromium	Cr	24	52.00
Cobalt	Co	27	58.93
Copper	Cu	29	63.55
Curium	Cm	96	245*
Damstadtium	Ds	110	271
Dubnium	Db	105	262
Dysprosium	Dy	66	162.5
Einsteinium	Es	99	254*
Erbium	Er	68	167.3
Europium	Eu	63	152.0
Fermium	Fm	100	254*
Fluorine	F	9	19.0
Francium	Fr	87	223*
Gadolinium	Gd	64	157.3
Gallium	Ga	31	69.72
Germanium	Ge	32	72.59
Gold	Au	79	197.0
Hafnium	Hf	72	178.5
Hassium	Hs	108	277
Helium	He	2	4.003
Holmium	Ho	67	164.9
Hydrogen	H	1	1.008
Indium	In	49	114.8
Iodine	I	53	126.9
Iridium	Ir	77	192.2
Iron	Fe	26	55.85
Krypton	Kr	36	83.80
Lanthanum	La	57	138.9
Lawrencium	Lw	103	260*
Lead	Pb	82	207.2
Lithium	Li	3	6.941
Lutetium	Lu	71	175.0
Magnesium	Mg	12	24.31
Manganese	Mn	25	54.94
Meitnerium	Mt	109	268
Mendelevium	Md	101	258*
Mercury	Hg	80	200.6
Molybdenum	Mo	42	95.94
Neodymium	Nd	60	144.2
Neon	Ne	10	20.18
Neptunium	Np	93	237.0
Nickel	Ni	28	58.71
Niobium	Nb	41	92.91
Nitrogen	N	7	14.01
Nobelium	No	102	255*
Osmium	Os	76	190.2
Oxygen	O	8	16.00
Palladium	Pd	46	105.4
Phosphorus	P	15	30.97
Platinum	Pt	78	195.1
Plutonium	Pu	94	244*
Polonium	Po	84	209*
Potassium	K	19	39.102
Praseodymium	Pr	59	140.9
Promethium	Pm	61	145*
Protactinium	Pa	91	231.0
Radium	Ra	88	226.0
Radon	Rn	86	222*
Rhenium	Re	75	186.2
Rhodium	Rh	45	102.9
Roentgenium	Rg	111	272
Rubidium	Rb	37	85.47
Rutherfordium	Rf	104	261
Ruthenium	Ru	44	101.1
Samarium	Sm	62	150.4
Scandium	Sc	21	44.96
Seaborgium	Sg	106	271
Selenium	Se	34	78.96
Silicon	Si	14	28.09
Silver	Ag	47	107.9
Sodium	Na	11	22.99
Strontium	Sr	38	87.62
Sulphur	S	16	32.064
Tantalum	Ta	73	180.9
Technetium	Tc	43	99*
Tellurium	Te	52	127.6
Terbium	Tb	65	158.9
Thallium	Tl	81	204.4
Thorium	Th	90	232.0
Thulium	Tm	69	168.9
Tin	Sn	50	118.7
Titanium	Ti	22	47.90
Tungsten	W	74	183.9
Uranium	U	92	238.0
Vanadium	V	23	50.94
Xenon	Xe	54	131.3
Ytterbium	Yb	70	173.0
Yttrium	Y	39	88.91
Zinc	Zn	30	65.38
Zirconium	Zr	40	91.22

*** Mass number of most stable isotope Source of Information*:*** Atomic weights of the elements 2005 (IUPAC Technical Report) (*Pure Appl. Chem., 2006, Vol. 78, No. 11, pp. 2051-2066*)

Appendix 2: The Periodic Table of Elements

Group

Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** hydrogen [1.007, 1.009]																	2 **He** helium 4.003
2	3 **Li** lithium [6.938, 6.997]	4 **Be** beryllium 9.012											5 **B** boron [10.80, 10.83]	6 **C** carbon [12.00, 12.02]	7 **N** nitrogen [14.00, 14.01]	8 **O** oxygen [15.99, 16.00]	9 **F** fluorine 19.00	10 **Ne** neon 20.18
3	11 **Na** sodium 22.99	12 **Mg** magnesium [24.30, 24.31]											13 **Al** aluminium 26.98	14 **Si** silicon [28.08, 28.09]	15 **P** phosphorus 30.97	16 **S** sulfur [32.05, 32.08]	17 **Cl** chlorine [35.44, 35.46]	18 **Ar** argon 39.95
4	19 **K** potassium 39.10	20 **Ca** calcium 40.08	21 **Sc** scandium 44.96	22 **Ti** titanium 47.87	23 **V** vanadium 50.94	24 **Cr** chromium 52.00	25 **Mn** manganese 54.94	26 **Fe** iron 55.85	27 **Co** cobalt 58.93	28 **Ni** nickel 58.69	29 **Cu** copper 63.55	30 **Zn** zinc 65.38(2)	31 **Ga** gallium 69.72	32 **Ge** germanium 72.63	33 **As** arsenic 74.92	34 **Se** selenium 78.97	35 **Br** bromine [79.90, 79.91]	36 **Kr** krypton 83.80
5	37 **Rb** rubidium 85.47	38 **Sr** strontium 87.62	39 **Y** yttrium 88.91	40 **Zr** zirconium 91.22	41 **Nb** niobium 92.91	42 **Mo** molybdenum 95.95	43 **Tc** technetium	44 **Ru** ruthenium 101.1	45 **Rh** rhodium 102.9	46 **Pd** palladium 106.4	47 **Ag** silver 107.9	48 **Cd** cadmium 112.4	49 **In** indium 114.8	50 **Sn** tin 118.7	51 **Sb** antimony 121.8	52 **Te** tellurium 127.6	53 **I** iodine 126.9	54 **Xe** xenon 131.3
6	55 **Cs** caesium 132.9	56 **Ba** barium 137.3	57-71 lanthanoids	72 **Hf** hafnium 178.5	73 **Ta** tantalum 180.9	74 **W** tungsten 183.8	75 **Re** rhenium 186.2	76 **Os** osmium 190.2	77 **Ir** iridium 192.2	78 **Pt** platinum 195.1	79 **Au** gold 197.0	80 **Hg** mercury 200.6	81 **Tl** thallium [204.3, 204.4]	82 **Pb** lead 207.2	83 **Bi** bismuth 209.0	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
7	87 **Fr** francium	88 **Ra** radium	89-103 actinoids	104 **Rf** rutherfordium	105 **Db** dubnium	106 **Sg** seaborgium	107 **Bh** bohrium	108 **Hs** hassium	109 **Mt** meitnerium	110 **Ds** darmstadtium	111 **Rg** roentgenium	112 **Cn** copernicium	113 **Nh** nihonium	114 **Fl** flerovium	115 **Mc** moscovium	116 **Lv** livermorium	117 **Ts** tennessine	118 **Og** oganesson

Lanthanide Series	57 **La** lanthanum 138.9	58 **Ce** cerium 140.1	59 **Pr** praseodymium 140.9	60 **Nd** neodymium 144.2	61 **Pm** promethium	62 **Sm** samarium 150.4	63 **Eu** europium 152.0	64 **Gd** gadolinium 157.3	65 **Tb** terbium 158.9	66 **Dy** dysprosium 162.5	67 **Ho** holmium 164.9	68 **Er** erbium 167.3	69 **Tm** thulium 168.9	70 **Yb** ytterbium 173.0	71 **Lu** lutetium 175.0
Actinide Series	89 **Ac** actinium	90 **Th** thorium 232.0	91 **Pa** protactinium 231.0	92 **U** uranium 238.0	93 **Np** neptunium	94 **Pu** plutonium	95 **Am** americium	96 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Appendix 3: SI prefixes, symbols and values

The following table, shows SI prefixes, their symbols and their values, and is useful when answering questions that involve the conversion of units:

SI Prefix	tera	giga	mega	kilo	centi	milli	micro	nano	pico
Symbol	T	G	M	k	c	m	μ	n	p
Value	10^{12}	10^{9}	10^{6}	10^{3}	10^{-2}	10^{-3}	10^{-6}	10^{-9}	10^{-12}